ELEMENTARY CHEMISTRY

FOR HIGH SCHOOLS AND ACADEMIES

BY

ALBERT L. AREY

ROCHESTER (N.Y.) HIGH SCHOOL

☞ *This pamphlet is sold at the retail price of Arey's "Chemistry," and can be exchanged at any bookseller's about November 1, without further charge, for a copy of the complete work.*

New York

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CHAPTER I

CHEMICAL ACTION

1. Elements and Compounds. — The different kinds of matter known to man may be divided into two classes:

(1) *Compounds*, or those which may be decomposed or separated into other substances.

(2) *Elements*, or those which have thus far resisted all attempts to decompose them.

About seventy simple substances or elements have been discovered, and, so far as is at present known, all compounds are the result of the chemical union of two or more of these.

It is possible that, as our knowledge of chemistry increases, many, if not all, of the substances now classed as elements may be shown to be compounds. Water was considered an element until 1783, and several other so-called elements have been resolved into simpler forms since that time. There are many chemists who consider the seventy elements as so many unsolved problems.

2. Molecules and Atoms. — The accepted theory of the constitution of matter maintains: —

1. That it is made up of minute particles called molecules (little masses), each one of which, in a given substance, is exactly like its neighbors in *weight*, *volume*, and *structure*.

2. That they move about each other, under the influence of heat, as separate bodies.

3. That they are the limit beyond which it is impossible to subdivide matter without destroying its identity.

In accordance with this theory, each molecule of a compound is believed to contain the same elements that chemical analysis shows the large masses of the substance to contain, and these smaller portions of the elements are called atoms. There is good reason for believing that atoms rarely exist in a free state, but that the molecules of most elements consist of two or more atoms.

A molecule is the smallest particle of a substance which can exist in the free state, and which has the same composition as any larger mass of the same substance.

An atom is the smallest particle of an element that exists in any molecule.

We may now state the following definitions: —

A compound is a substance whose molecule contains two or more kinds of atoms.

An element is a substance whose molecule contains only one kind of atoms.

3. The Domain of Chemistry. — *Chemistry is that branch of science which deals with changes in the identity of substances; and with the laws, causes, and effects of such changes.*

The subject is closely related to physics; every chemical change is accompanied by some physical change, but the chemical change differs in one important particular from a physical change: the chemical change is due to forces acting upon atoms, while the physical change depends upon forces acting upon the molecule.

A physical change is one which does not destroy the identity of the substance acted upon.

Illustration. When a bar of steel is magnetized it acquires a new property, but it remains the same substance, and the change is physical.

A chemical change is one which destroys the identity of the substance acted upon.

Illustration. When a bar of steel rusts a portion of the steel is converted into a new substance which differs from the steel in color, tenacity, elasticity, and other properties.

4. Chemical Action. — In some instances one may be in doubt as to whether a chemical change has taken place, and in a few instances chemical analysis is necessary to prove it. But in general the occurrence of any of the following phenomena, when two or more substances are mixed, may be taken as evidence of chemical action: —

1. Effervescence.
2. The evolution of heat and light.
3. Change of color.
4. Change of volume.
5. Change of state.
6. The development of electricity.

Exceptions. 1. A change of state by solution of a solid or gas.

2. A change of volume due to the absorption of a gas by a solid or liquid, or to a change in temperature.

Take notes on the following experiments which will be performed for you, and designate them as physical or chemical changes: —

Experiment I. — Sugar and potassium chlorate are mixed, and a drop of sulfuric acid added.

Experiment II. — Sulfuric acid is added to syrup.

Experiment III. — A rubber ruler is electrified.

Experiment IV. — A beam of sunlight is decomposed with a prism.

Experiment V. — A piece of platinum wire is heated to redness.

Experiment VI. — A piece of magnesium wire is heated to redness.

Experiment VII. — Sulfur and potassium chlorate are mixed in a mortar with considerable friction.

Experiment VIII. — Solutions of potassium iodid and mercuric chlorid are mixed.

Experiment 9 and the succeeding experiments are to be performed by the pupil unless special directions to the contrary are given.

Experiment IX. A Chemical Change. — 1. Examine a piece of marble carefully, fix its appearance in mind, so that you can detect any change.

2. Drop a small piece in the test bottle and cover it with dilute hydrochloric acid. What occurs?

3. After a short time test the gas in the upper part of the test bottle with a lighted match. Does the match continue to burn? Is the gas combustible? Is the gas ordinary air? Why? Does the marble disappear?

4. In order to tell whether the marble has been changed chemically, the acid must be expelled. To accomplish this, pour the solution into an evaporating dish, place it on a piece of wire gauze, and bring the liquid to a boil. When the liquid begins to solidify and turn yellow, add a few drops of water, repeating, if necessary, to obtain a solid white residue.

Examine this residue, compare it carefully with marble. Set the residue aside for 24 hours to determine whether it is permanent when exposed to the air.

Fill out the following table: —

	Marble	Residue
Is it hard or soft?		
Does it effervesce with hydrochloric acid?		
Is it soluble in water?		
Is it permanent in air?		

How do the properties of the residue compare with those of marble?
Is it marble?
What have you proven?

Experiment X. — Bring together on a flower-pot saucer a little phosphorus and iodin. What evidence have you that chemical action took place? Have either of the original substances disappeared? Has a new substance been formed? It will be seen that simple contact is sufficient to cause the two substances to act upon each other.

Does either substance melt? Why?

Is this a case of chemical action between solids?

Is the action as energetic at first as it is after a few seconds? Explain.

CAUTION. — Handle phosphorus with great care; it takes fire when rubbed or cut in the air, and should always be kept in water.

5. Conservation of Matter.

Experiment XI. — Pour 10 cc. of dilute sulfuric acid into a beaker. In a second beaker pour an equal quantity of calcium chlorid solution. Place both beakers in one scale pan and balance them carefully with weights, sand, or shot, placed in the other scale pan. Now pour the calcium chlorid into the sulfuric acid. Does a chemical change occur? Replace the beakers and determine whether the weight of the beakers and their contents has been changed. Does chemical action change the total quantity of matter in existence? Was the total quantity of sulfuric acid in the world increased or diminished by the above experiment? How was the total quantity of calcium chlorid affected? of the white substance formed?

6. The Effect of Solution on Chemical Action.

Experiment XII. — Place as much baking soda as you can take on the end of a knife blade in a dry test bottle. Add an equal amount of tartaric acid; shake the bottle to mix the powders thoroughly. Has any change occurred?

Pour a few cubic centimetres of water into the bottle. What evidence of chemical action do you observe?

Does solution aid chemical action? Is it because more intimate contact of the molecules is obtained when solutions are mixed than is possible with solids? Should diminishing cohesion assist chemical action? State your opinion as to why solution aids chemical action.

7. Effect of Heat on Chemical Action.

Experiment XIII. — 1. Mix six grammes of potassium chlorate and one gramme of powdered charcoal thoroughly. What occurs?

2. Apply a lighted match. Was the change chemical or physical?

How does the operation of striking a match illustrate the effect of heat upon chemical action?

Why do metals rust more rapidly when hot than at lower temperatures? Experiment 22 illustrates this effect. Do fuels combine with the air when cold?

8. Light causes Chemical Action.

Experiment XIV. — Cut a design from tin-foil and place it on a piece of blue print paper. Expose paper and design to sunlight for a few minutes. Wash the paper in water.

Has the sunlight affected the exposed chemical? In what way?

The art of photography is based on the action of light on chemicals. In growing plants sunlight causes the decomposition of carbon dioxid, which is only accomplished by the chemist with difficulty.

In the preparation of hydrochloric acid by synthesis described on page 92, the chemical action is assisted by light.

QUERY. — Why do certain colors fade when exposed to light?

9. Pressure. — When the two gases, hydrochloric acid and hydrogen phosphid, are subjected to increasing pressure they combine to form a crystalline solid known as phosphonium chlorid. Similarly sulfur and powdered lead may be caused to combine by great pressure, forming lead sulfid.

QUERY. — What relation does this action suggest between the intensities of chemical affinity and distances between molecules?

10. Concussion or Detonation. — In a very few cases, chemical action is brought about by detonation. The molecules of the gas acetylene consist of two atoms of carbon united with two of hydrogen. If a small quantity of mercury fulminate be detonated near a globe filled with this gas the carbon is instantly deposited in solid form and the hydrogen liberated. This action is not fully understood; some chemists believe that the particular form of sound vibration produced disturbs the motions of the atoms constituting the molecule, and thus causes disruption.

11. Electricity. — If a current of electricity be passed through a solution of copper sulfate, the compound is de-

composed, and many other compounds are affected in the same way. In Experiment 40 this effect is also illustrated.

12. The Effect of Trituration on Chemical Action.

Experiment XV. — Using pincers, hold a small lump of rosin in the Bunsen burner flame, observe the character of the flame produced by the burning rosin. Is rosin easily ignited? Does it burn rapidly? Does the rosin melt before it ignites?

Experiment XVI. — Triturate a small piece of rosin in a mortar, fill the end of a large glass tube with the powder and blow it into the burner flame. Does the finely divided rosin burn with a smoky flame or does it flash? Does it burn as rapidly as in the previous experiment? How does the energy of the chemical action compare with that observed in the last experiment? In which case is the higher temperature reached?

Experiment XVII. — Make a compact pile of about $\frac{1}{2}$ cu. cm. of powdered rosin on a piece of porcelain or earthenware, ignite with a Bunsen burner. How does the chemical action compare with that of the previous experiment? Does the increased chemical action depend upon the size of the particles? Would a solid piece having the same area as the sum of the surfaces flash? Does the chemical activity depend upon the surface only? Upon the mass of the particles only?

13. Mechanical Mixture.

Experiment XVIII. — 1. Mix about four grammes of sulfur and an equal weight of fine wrought iron filings on a sheet of paper. Divide into three portions.

2. Examine the first portion with a magnifying glass. Can you distinguish the particles of sulfur from those of iron? Can you separate the iron from the sulfur with a magnet? Now put the mixture in a test tube and pour water on it. Are the substances combined or not? Shake the tube; what is the yellow substance floating on the water? Has chemical action taken place?

3. Treat the second portion with carbon disulfid. What is the black substance at the bottom of the tube? What has happened? Is the color of the carbon disulfid changed? What does this indicate? Is a chemical compound formed in this experiment?

Experiment XIX. — 1. Put the third portion of the mixture made in Experiment 18, in a dry test tube and heat gently. When it is red hot remove the tube from the flame. Is there any evidence of combustion in the tube?

2. After the action is over and the tube has cooled down, loosen the contents with a short piece of wire, and pour it out on a piece of paper. Does the mass look like the mixture of sulfur and iron with which you started?

3. Examine with a magnifying glass. Can you separate the sulfur and iron with water as before? Can you separate them with a magnet?

4. Treat a portion of the mass with carbon disulfid. Is the effect the same as before? Is the color of the carbon disulfid changed? What do you conclude concerning the effect of heat on the mixture?

REVIEW QUESTIONS

1. Define chemical action. What assists it? What retards it?

2. Mention those conditions which aid chemical action, (*a*) by decreasing the distance between the unlike molecules, (*b*) by diminishing the cohesion of the factors.

3. Describe an experiment to show that there is no loss of matter in chemical change.

4. Distinguish between a mechanical mixture and a chemical compound. Illustrate each.

5. Distinguish between chemistry and physics; between atoms and molecules; between chemical changes and physical changes.

6. What mechanical mixture was formed in Experiment 12? In what part of the experiment were chemical compounds formed?

Write answers to these questions in your note-book.

CHAPTER II

SYMBOLS AND LAWS

14. Symbols. — Chemists of all countries have agreed to use the initial letter of the Latin name of an element as an abbreviation which shall stand for a single atom of that element. In case two or more elements begin with the same letter the second characteristic letter is added to the symbol, thus: —

C	Carbon	N	Nitrogen	Na *	Sodium
Ca	Calcium	S	Sulfur	K *	Potassium
Cl	Chlorin	Si	Silicon	Ag	Silver

Some writers use these symbols as mere shorthand signs for the full names of the elements. This usage is extremely objectionable; students who adopt it will not appreciate the important quantitative relations which are shown by reactions.

15. Formulæ. — Compounds are represented by a formula or a group of symbols, showing the composition of the molecule of the substance.

Thus, the formula of sodium chlorid, NaCl, indicates that its molecule contains one atom of sodium and one of chlorin, and CaS represents a molecule of calcium sulfid which contains one atom of calcium and one of sulfur.

If a molecule contains more than one atom of a given

* The Latin name of sodium is *Natrium*, that of potassium is *Kalium*.

element, a subnumber is placed a little below and to the right of the symbol, and indicates the number of such atoms. Thus $CaCl_2$ is the formula for calcium chlorid, which contains one atom of calcium and two of chlorin, and the formula for ferric oxid, Fe_2O_3, tells us that its molecule contains two atoms of iron and three of oxygen.

If more than one molecule of the substance is to be represented, the number is placed before the group of symbols. Thus $2\,Fe_2O_3$ represents two molecules of ferric oxid containing four atoms of iron and six of oxygen.

In the absence of a coefficient a formula always represents a single molecule.

16. The Law of Constant Proportions. — The law of definite proportions which has been called the corner stone of modern chemistry is as follows: —

"*The same compound always contains the same elements combined in the same fixed and definite proportions.*"

The thousands of analyses which have been made of various compounds by chemists in all parts of the world, and which are now being made every day, are based upon this law, and in no single instance have the results obtained caused the truth of the law to be questioned.

17. Combining Weights. — Another important relation is to be learned from a study of the composition of various substances. Not only is the proportion by weight in which a certain element combines with a certain other element, to form a given compound, constant, but it is possible to select a number for each element, which shall represent the proportion by weight in which it unites with different elements.

The composition of the oxids mentioned thus far is given below: —

Mercury,	200	Lead,	207	Iron,	56
Oxygen,	16	Oxygen,	16	Oxygen,	16
Copper,	63.6	Zinc,	65.4		
Oxygen,	16	Oxygen,	16		

In each of the above compounds it is observed that there are 16 parts by weight of oxygen, and this number or a simple multiple of it will express the proportion in which oxygen combines with any other element.

Such numbers have been carefully determined for all elements and are called combining weights.

18. The Law of Multiple Proportions. — The analysis of various substances further shows that a given element may combine with another in more than one proportion. For example, the elements nitrogen and oxygen form several compounds having the following composition: —

	Nitrogen	Oxygen
Nitrous oxid	28 parts	16 parts
Nitric oxid	28 parts	32 parts
Nitrous anhydrid	28 parts	48 parts
Nitrogen peroxid	28 parts	64 parts
Nitric anhydrid	28 parts	80 parts

It will be observed that while the quantity of nitrogen is the same in all the above compounds the quantity of oxygen varies, being twice as great in the second compound as in the first, three times as great in the third as in the first, etc. This series illustrates the law which applies to all cases in which more than one compound is formed from the same elements.

If two elements form more than one compound, the proportions by weight in which a given element combines with the other in each compound will be expressed either by its combining number or a simple multiple of its combining number.

19. The Atomic Theory. — The atomic theory was suggested by John Dalton, an English schoolmaster, early in this century, to account for the laws of definite and multiple proportions. It maintains —

1. That with a few possible exceptions all molecules are made up of smaller particles.
2. That these particles are indivisible (they are therefore called atoms).
3. That all atoms of a given element are equal in size and weight.
4. That atoms of different substances have different weights.
5. That the combining weights of the elements are simply the relative weights of the atoms, and may therefore be called the atomic weights.

The explanation of the facts of chemistry which this theory offers is so satisfactory that it is universally accepted.

20. Atomic Weights. — As hydrogen enters into combination in smaller proportion than any other element, its combining weight or atomic weight is taken as the unit. When we say that the atomic weight of oxygen is 16, we mean simply that the atoms of oxygen are sixteen times heavier than those of hydrogen. The exact weight of an atom of hydrogen has never been determined but it is called a *microcrith.* The atom of oxygen weighs 16 microcriths.

The following table gives the exact values of the atomic weights of the elements referred to in this book. The standard is Oxygen = 16.

TABLE OF ATOMIC WEIGHTS

Name	Sym.	O = 16	Name	Sym.	O = 16
Aluminum . . .	Al	27.11	Antimony . . .	Sb	120.42
Argon	A	(?)	Arsenic . . .	As	75.01
Barium	Ba	137.43	Bismuth . . .	Bi	208.11
Boron	B	10.95	Bromin	Br	79.95
Cadmium . . .	Cd	111.95	Calcium . . .	Ca	40.07
Carbon	C	12.01	Chromium . . .	Cr	52.14
Chlorin	Cl	35.45	Copper	Cu	63.60
Cobalt	Co	58.93	Gold	Au	197.23
Fluorin	F	19.06	Iodin	I	126.85
Hydrogen . . .	H	1.008	Lead	Pb	206.92
Iron	Fe	56.02	Magnesium . .	Mg	24.28
Lithium	Li	7.03	Mercury . . .	Hg	200.00
Manganese . . .	Mn	54.99	Nitrogen . . .	N	14.04
Nickel	Ni	58.69	Phosphorus . .	P	31.02
Oxygen	O	16.00	Potassium . . .	K	39.11
Platinum . . .	Pt	194.89	Silver	Ag	107.92
Silicon	Si	28.40	Strontium . . .	Sr	87.61
Sodium	Na	23.05	Tin	Sn	119.0
Sulfur	S	32.0	Zinc	Zn	65.41

21. Reaction. — The force which is exerted between atoms is called *chemical affinity*. The affinity of a given atom for other atoms varies greatly, often being very strong for certain kinds of atoms and feeble for others. If, when any substances are mixed, a rearrangement of the atoms would produce more stable compounds, *i.e.* if the force which holds the atoms together in the new compounds is stronger than that which bound them in their original form, such rearrangement will take place. The process of redistribution of the atoms in the molecules concerned in the phenomenon is called *chemical action* or *reaction*.

A reaction is due to chemical affinity and causes a chemical change.

Substances used to bring about desired reactions are called *reagents.*

The substances which go into a reaction are called *factors*, and those which come from a reaction *products.*

Reactions are ordinarily expressed by *equations* in which the symbols and formulæ of the factors are placed on the left of the sign of equality, and those of the products on the right. The algebraic signs plus and minus are used in the ordinary sense in the equations.

The fact that atoms can neither be created nor destroyed, even by chemical means, justifies the use of the sign of equality to connect factors and products, and it should never be placed until the student has "satisfied the reaction," *i.e.* has determined that there are exactly as many atoms of each element in the products as there are in the factors.

Illustration. In Experiment 11 the following reaction occurred: —

$$CaCl_2 + H_2SO_4 = CaSO_4 + 2\,HCl.$$

This should be read as follows: one molecule of calcium chlorid plus one molecule of sulfuric acid forms one molecule of calcium sulfate plus two molecules of hydrochloric acid.

The chemical change occurring in Experiment 19 may be expressed as follows: —

$$Fe + S = FeS.$$

Such equations express very concisely the relations between the atoms and molecules in the chemical changes which they represent, and every chemical change which is clearly understood may be expressed in this way. Equations are also useful because of the important quantitative relations between masses which are made evident when we consider the atomic weights of the elements represented.

In the equation given above, the symbol Fe not only signifies an atom of iron, but it also stands for 56 parts of iron, by weight, and the symbol S stands for 32 parts of sulfur by weight. We thus have a mathematical expression which shows the relation between the masses of iron and sulfur which take part in the chemical change.

$$\begin{array}{ccc} Fe + & S = & FeS \\ 56 & 32 & 88 \end{array}$$

The equation can now be read: —

56 parts of iron unite with 32 parts of sulfur to form 88 parts of ferrous sulfid. The solution of many chemical problems depends upon this use of equations. (See Chapter IX.)

22. Analysis, Synthesis, and Metathesis. — All chemical changes may be referred to one of four classes: —

(*a*) Compound molecules may be separated into their elements, or into simpler groups of elements, as, for example, mercury rust is separated into mercury and oxygen in Experiment 25, or as potassium chlorate $KClO_3$, is decomposed in Experiment 29, forming potassium chlorid KCl, and oxygen. Such changes are analytic, and the process which brings them about is known as *analysis*.

(*b*) Compound molecules may be formed by direct union of elements, or simpler groups of elements, as when phosphorus combined with iodin, in Experiment 10, forming phosphorous di-iodid PI_2, or when carbon monoxid CO, combines with oxygen to form carbon dioxid CO_2. (See Experiment 96.) Such changes are synthetic, and the process is known as *synthesis*.

(*c*) Compound molecules may be formed by a change involving both analysis and synthesis, which is known as *metathesis*, or double decomposition. In such processes an exchange of atoms, or groups of atoms, takes place between two compound molecules, as when a solution of sodium sulfate Na_2SO_4, and barium chlorid $BaCl_2$, are mixed. Each substance is decomposed, and the atoms combine to form two new substances, barium sulfate $BaSO_4$, and sodium chlorid NaCl.

(*d*) In some cases new substances are formed without changing either the kinds of atoms, or the number of atoms of each kind, in the molecule. For example, when a solution of ammonium cyanate $NH_4 . O . CN$, is heated it is transformed into urea N_2H_4CO, a substance having entirely different chemical and physical properties. It will be observed that these molecules contain the same number of atoms of each element; we have excellent evidence, however, that the first one contains cyanogen CN, while the second contains carbon monoxid CO.

REVIEW QUESTIONS

1. How many atoms of hydrogen in $6\ H_2SO_4$? of sulfur? of oxygen?

2. How many atoms of each element are represented by the following formulæ: $2\ ZnCl_2$, $3\ HNO_3$, $5\ H_2O$, $14\ NH_3$.

3. How many molecules of each substance are represented by above formulæ?

4. Define chemical affinity, reaction, reagent, factor, product.

5. Distinguish between atoms and molecules. Does a chemical affinity exist between molecules? Give a reason for your answer.

6. What is atomic weight? How is atomic weight related to specific gravity?

7. What element is selected as the standard of atomic weight? Why is this element selected?

8. State the atomic theory.

9. State five principles observed in writing chemical symbols and formulae.

10. What is a chemical equation? What is meant by the combining weight of an element?

11. State the law of constant proportions. Of multiple proportions.

12. State five principles to be observed in writing chemical equations.

CHAPTER III

CHEMISTRY OF THE AIR

23. The Formation of Rust.

Experiment XX. — 1. In a small porcelain crucible or a clay pipe bowl put a small piece of lead or zinc. Heat with laboratory burner and notice the changes that take place. Do not allow the containing vessel to become too hot, for liquefied rust will be absorbed. After the lead begins to melt, stir with a thick iron wire. Observe carefully what forms on the surface of the metal. Does the lead retain its bright mirror-like surface if not stirred? Continue to heat and stir until the substance is changed to a powder. What is its appearance now?

2. Let it cool. Is it lead? What difference is there between the action in this case and in melting ice and cooling the water again? Which is chemical and which is physical action? Why? Was the change just observed produced by the heat or by the action of the air? In order to answer this question let us repeat the experiment, preventing any action of the air by covering the metal with a film of melted rosin.

Experiment XXI. — Repeat Experiment 20, adding as much powdered rosin as can be lifted on the blade of a penknife. Do not stir the metal. Does it rust or does the surface remain bright and mirror-like? Is it changed to powder? How do you explain the difference in result of this experiment and the last? What do you conclude concerning the cause of the change produced in the previous experiment? Is the action due to the high temperature or to the action of the air or to both? Does lead rust more rapidly at high than at low temperature? Rosin is used to prevent rusting of hot metals in process of soldering.

24. Effects of Air on Iron at Ordinary and at High Temperatures.

Experiment XXII. — Wind a piece of No. 30 iron wire about a foot long around the finger and heat the loops thus formed in the tip of a laboratory burner flame for a minute or two. Holding the loop over a sheet of paper, straighten the wire. Compare the scale which drops

off with the rust formed on iron at ordinary temperatures. Is it the same color? Does iron rust more rapidly at high or low temperatures? How do you know? Has a chemical change occurred? Pass a magnet over a mass of red rust; of black rust. Are they magnetic? (See paragraph on Oxids in Nature, page 35.)

25. Various Ways of Protecting Iron.

Several years ago Professor Barff, of London, suggested that iron might be protected from the action of the air by exposing it to superheated steam at high temperature, thus forming a coating of black rust on its surface. The process has been somewhat modified, and is now known as the Bower-Barff process. It is quite extensively employed as a finish for iron ornaments, and has been used in certain cities to protect water pipes.

Zinc and lead are protected from the action of the air by the coating of oxid which forms on their surface.

QUERIES. — Mention several ways of protecting iron from the action of the air. Why do we blacken stoves? Why are some parts nickel plated? What is galvanized iron? What is a tin pan made of? In what two ways are water pipes protected? How are iron bridges protected? Bicycle frames?

26. Does the Weight of a Metal change when it rusts?

When a chemical change occurs it is due to the addition of some element or elements to the substance changed, or to the extraction of some element or elements from the substance changed. Now, since loss or gain in matter means loss or gain in weight, let us determine whether a substance was added to or driven off from the iron in the last experiment.

Experiment XXIII. (Performed by the instructor.) — Weigh a piece of No. 30 wire, heat as in the last experiment; when cool weigh again. Explain.

Does heating in contact with air drive something away from the iron or cause something to combine with it? From what source is the substance derived?

27. The Material which combines with the Metal to form Rust.

We now desire to know the nature of the substance which causes metals to rust; and as it can be expelled easily from the rust which forms on mercury, we shall study that substance. As air is an invisible gas, special precautions must be taken to prevent its loss or mixture with other substances.

Experiment XXIV A Method of Collecting Gases. — Fill the yellow dish (see description of apparatus, p. 000) one-third full of water. Place the flower-pot saucer bottom side up in the water. Fill one of the medium sized bottles with water, cover with a glass plate, and invert on the flower-pot saucer; remove the glass plate. If your work has been carefully performed your bottle will be full of water. (If not, try again.)

Now put the end of a glass tube at the opening at the side of the flower-pot saucer and blow gently through it. What do you notice? What is in the bottle after the water is out of it? Where does it come from?

This method of collecting gases over water may be used for all gases not dissolved by water.

Students should attempt to devise other methods. Could a rubber bag be used? What advantage has the method used in this experiment over other methods?

Experiment XXV. A Study of Mercury Rust. — 1. Weigh accurately a small glass tube, closed at one end, containing about a gramme of mercury rust.

2. Holding the tube in a nearly horizontal position with a pair of crucible tongs, heat the red powder strongly for some minutes, or until a bright mirror-like deposit appears near the open end of the tube.

3 Weigh the tube again. Is there any evidence that an invisible substance has escaped? After weighing the tube, examine the deposit near the open end. Scrape some of it from the tube with an iron wire; what is it? What have you learned about the constituents of mercury rust? Is either constituent a solid? a liquid? a gas?

4. Arrange an ignition tube, as shown in Fig. 1, so that any gas generated in the tube may be collected in the bottle. Fill the bottle with water.

5. Put about 15 grammes of mercury rust in the ignition tube and apply heat. Describe the gas collected.

6. Test with a glowing match stick. Remove the match and put it back a few times. Is there any difference between the burning in the bottle and out of it? Is the gas air? Has the gas which formed the rust a marked ability to make things burn? Has the color of the mercury rust changed?

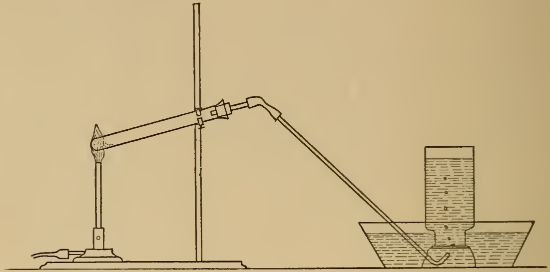

FIG. 1.

7. Remove the ignition tube and pour its contents on a piece of paper. How is the color affected? Compare it with some of the mercury rust which has not been heated. What effect has the air had upon the hot rust from the tube? Has the air entirely restored the gas driven off by the heat? Is the gas collected in this experiment pure air, or a part of the air?

The chemical change which occurs in this experiment may be expressed as follows: — $HgO = Hg + O$.

The gas which causes metals to rust is called *oxygen*, its compounds are called *oxids*, and the process of forming oxids is known as *oxidation*. The rust formed on iron at ordinary temperatures is called ferric hydroxid. That formed at high temperatures is ferrous oxid and ferric oxid, probably in chemical combination. It is called magnetic oxid.

We have observed that oxygen makes things burn vigorously, and, although it is deemed best to reserve the discussion of combustion for a subsequent chapter, the next two experiments are given here to show the relation between the processes of rusting and burning.

28. Effect of excluding the Air from a Flame.

Experiment XXVI. — Close the holes at the bottom of your laboratory burner. How is the character of the flame affected? Explain. Why do we close the stove dampers at night? What is the effect of removing the ashes and clinker from a stove? Why?

Suggestion. — Wrap a piece of cloth around the lower part of a kerosene lamp burner, covering the holes through which the air enters the chimney. How does this affect the flame? What has air to do with the combustion of oil? How does a lamp chimney increase the brightness of the lamp flame? How does a fire extinguisher put out a fire? It is possible that burning, like rusting, is simply a chemical union of air, or a part of the air, with the fuel. Let us determine whether this is so by the method used in Experiment 23.

29. Comparison of the Weight of the Products of a Burning Candle with the Amount lost by the Candle.

Experiment XXVII. (Performed by the instructor.)— On one side of a delicate balance, apparatus which will absorb the products of combustion is suspended over a candle, the whole being exactly balanced with weights on the other side of the balance. The candle is lighted and the gases are drawn into the absorbing apparatus. As the candle burns away the side of the balance carrying the apparatus grows heavier. The weight of the products is greater than the loss of weight of the candle.

Is the candle converted into heat? Is heat matter? Where does the matter causing the increase come from? Does this prove that the candle is indestructible? What is your conclusion concerning the nature of combustion?

30. Analysis of the Air. — We have learned that oxygen is a part of the air, and now desire to learn what proportion of the air is oxygen.

Experiment XXVIII. To determine the per cent of oxygen in the air. Cooley's method. Apparatus required. — A small glass funnel. A six-inch test tube, with a two-holed rubber stopper to fit same. Rubber bands, a measuring glass, 155 cc. of the absorbent liquid. A piece of glass tubing two inches long fitted in one of the holes of the stopper, a piece of glass rod the same length in the other hole, and a piece of thin rubber tubing six inches long, in which a piece of glass

rod half an inch long and of such size as to prevent a liquid from running through the tube, is placed.

Manipulation. — 1. Arrange the apparatus as shown in Fig. 2.

2. In your test bottle dissolve a small teaspoonful of pyrogallic acid in 10 cc. of water, quickly add 5 cc. of a strong solution of sodium hydroxid, and pour into the funnel. This liquid absorbs oxygen and carbon dioxid rapidly.

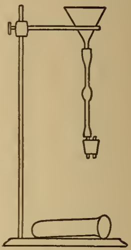

Fig. 2.

3. Holding the test bottle under the rubber cork, pinch the rubber tube where the glass rod closes it until a little of the liquid runs through the tube. Carefully remove the drop which is suspended from the glass tube with a piece of filter paper.

4. Now remove the glass rod from the hole in the rubber stopper and put the test tube on the stopper; allow it to hang there a minute or two to allow the heat communicated to the tube and air which it contains to pass away. Then insert the glass rod in the open hole in the stopper. We have now isolated a definite volume of air at the same temperature and pressure as the air of the room, and during the absorption and the measurements care must be taken to prevent change in the volume under analysis, either by the escape of a portion or by the introduction of more air from without.

5. Pinch the rubber tube at the glass rod to allow some of the absorbent liquid to run down into the test tube; a little stream runs in at first, then drops follow each other more and more slowly; when these have nearly ceased allow the apparatus to stand for two or three minutes. Then allow more of the absorbent liquid to enter the test tube. Repeat the operation every two minutes until only a drop or two enters the tube when opened.

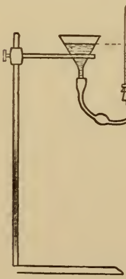

Fig. 3.

6. The gas in the test tube is now compressed by the weight of the liquid in the rubber tube; before measurements can be made the pressure must be adjusted to that of the air in the room. This is accomplished by grasping the test tube by the flange (so as not to warm the gas), raising the tube as shown in Fig. 3, and pinching the rubber tube to open a passage between the two masses of liquid. Keep this passage open and move the test tube up or down until the liquid stands at the same level in the test tube

and the funnel; then close the passage between them. Your results will depend to a great extent upon the care with which this adjustment is made.

7. Slip a rubber band around the test tube so that its upper edge marks the position of the bottom of the stopper.

8. Remove the test tube from the apparatus and pour the absorbent liquid into a measuring glass. This represents the volume of gas absorbed; record the number of cubic centimetres.

9. Now fill the test tube with water to the top of the rubber band and measure this volume. This represents the volume of air analyzed.

We have thus determined the number of cubic centimetres of oxygen in a certain number of cubic centimetres of air, from which we may determine the number of cubic centimetres of oxygen in 100 cc. of air; *i.e.* the percentage of oxygen in air.

31. Other Substances in the Air. — When a gas called carbon dioxid is passed through lime water the latter becomes cloudy because a white solid (a precipitate) is formed. This is the test for carbon dioxid.

Experiment XXIX. — 1. Take 20 or 30 cc. of lime water in your test bottle. Blow through a glass tube in such a way that the exhaled air bubbles through the lime water. Does the lime water become cloudy or does it remain clear? What does this experiment prove regarding air exhaled from the lungs?

2. Force air from a bellows through lime water. What inference do you draw from this experiment?

Carbon dioxid was absorbed with the oxygen in Experiment 28, but the amount is so small (about $\frac{1}{20}$ of one per cent) that it may be disregarded.

Does water vapor exist in the air? To answer this question, think of the moisture which collects on the outside of an ice pitcher in summer. What is dew? What is frost?

The gas which remains in the apparatus after absorbing the oxygen and carbon dioxid (Experiments 28 and 30) is nearly pure nitrogen. Nitrogen is fully discussed in Chapter VI.

ARGON

SYMBOL A. — ATOMIC WEIGHT 19.9

32. Some years ago Lord Rayleigh proved that nitrogen obtained by removing the oxygen from the air was invariably denser than that obtained from chemical compounds. He undertook to determine the cause of this difference, and in conjunction with Professor Ramsay found that this greater density was due to the presence of an unknown gas, which they succeeded in isolating and to which they gave the name Argon. Their discovery was announced January 27, 1895.

Argon is a gas forming $\frac{1}{120}$ part of the air; it is also found among the occluded* gases in some specimens of meteoric iron. As indicated by its name, argon is the most inert element; it has thus far resisted all attempts to get it to combine with other elements. Its chief characteristic, therefore, is its "glorious uselessness." It is sparingly soluble in water, boils at $-187°$ C. and freezes at $-189°$ C.

Since the discovery of argon several other new elements have been found in the air, with properties quite similar to those of argon.

33. Air as a Mixture. — Air is believed to be a mechanical mixture of nitrogen and oxygen, and not a chemical compound, for the following reasons: —

1. Air contains approximately 79 % of nitrogen and 21 % of oxygen. This is not in accordance with the law of multiple proportions.

2. If nitrogen and oxygen be mixed in the above proportions the mixture possesses all the properties of air, but is not accompanied by any phenomena which indicate chemi-

* Define term.

cal action. Whenever chemical union takes place, there is some change in the temperature of the substance; when nitrogen and oxygen are mixed as above described there is no change in the temperature.

3. The law of definite proportion states that the composition of a given chemical substance is invariable; that of air varies slightly.

4. Air is somewhat soluble in water, but each gas is dissolved independently.

If we shake up air and water in a bottle some of the air will be dissolved; if we boil this saturated water the air which escapes can be collected and analyzed. This has often been done, and it has been found to contain a larger proportion of oxygen than the original atmospheric air.

Thus	Atmospheric	Dissolved
N	79.04	66.36
O	20.96	33.64

This change in the proportion could not occur if the air was a compound, for a compound is dissolved as a whole. The above numbers exactly agree with the solubilities of oxygen and nitrogen separately.

REVIEW QUESTIONS

1. Describe the effects of the partial and of the total exclusion of air from a flame.

2. State how the effect of air on iron at high temperatures differs from the effect of air on iron at ordinary temperatures.

3. Describe a chemical method of protecting iron from the action of the air.

4. How does the weight of the products of the combustion compare with the amount lost by the candle? Why?

5. Does the weight of the scale which flies from the blacksmith's hot iron equal the weight lost by the iron? Why?

6. Has air a chemical formula?

7. Is the air a mixture or a compound? Describe an experiment to prove the correctness of your answer. Give reasons for your answer.

8. Describe the Bower-Barff process of protecting iron.

9. Give an account of the discovery of argon. State the properties of argon and its occurrence in nature.

10. Explain the effect of excluding air from a flame. Mention some practical appliance whose efficiency depends on the principle involved.

11. What is the scale which accumulates about the blacksmith's anvil?

CHAPTER IV

OXYGEN

Symbol O. — Atomic Weight 16

34. Occurrence. — Oxygen is the most abundant of all the elements, comprising by weight $\frac{1}{5}$ of the air, $\frac{8}{9}$ of the water, $\frac{3}{4}$ of all the animal bodies, and about $\frac{1}{2}$ of the crust of the earth.

The word oxygen means "acid-former," but it is a misnomer. Chemists supposed that it was present in all acids when the name was given.

35. Preparation. — Oxygen may be easily obtained by heating potassium chlorate.

Caution. — The following precautions must be observed : —

1. The chemicals must be free from impurities which might cause an explosion. If a small quantity of the mixture when heated in a dry test tube melts quietly, the mixture may be considered safe.
2. The ignition tube must be inclined.
3. It must not be more than one-third full.
4. The upper part of the mixture in the tube should be heated first.
5. The heat must be so regulated that an even and not too rapid flow of the gas may be secured. It may be necessary to withdraw the flame and replace it when the gas slackens.

Experiment XXX. (Two students will work together.) — 1. Arrange the apparatus as in Experiment 25. Mix equal weights of manganese dioxid and potassium chlorate, and heat about ten grammes of the mixture in a test tube. Collect four bottles of the gas evolved over water.

2. Place the bottles on the table, mouth upwards, covering them with a glass plate. What is the color of the gas ? Odor ? Taste ? Is it soluble in water ? The slight cloud which appears in the bottle

at first is due to a substance which is not oxygen. After a while this disappears and oxygen remains.

3. Drop a piece of charcoal, obtained by charring the end of a match stick, in the first bottle. In another lower a deflagrating spoon containing a little sulfur.

4. In the third drop a piece of phosphorus about the size of a pea. (Care!) Let them stand quietly and observe what changes, if any, take place. Does oxygen at ordinary temperatures act readily on these substances?

5. Now thrust a piece of red-hot charcoal (a glowing match stick) into the first bottle. Note difference in action.

6. Remove the deflagrating spoon from the second bottle; set fire to the sulfur. Notice whether it burns with ease or with difficulty. Does the sulfur burn more readily in the oxygen than in the air?

7. Remove the phosphorus from the third bottle; place it in the deflagrating spoon, ignite, and quickly lower it into the bottle again. Describe the action. How does the action of oxygen on these substances at high temperatures compare with the action on the same substances when cold? Does either substance burn as vigorously in air as in oxygen?

Reaction: $2\,KClO_3 + MnO_2 = 2\,KCl + MnO_2 + 3\,O_2$.

The Test for Oxygen. — Thrust a glowing splinter of wood into one of the bottles. What occurs?

NOTE. — No substance but oxygen can cause a spark to burst into flame. How can you determine whether a bottle contains oxygen or not?

36. Physical Properties. — Pure oxygen is colorless, odorless, and tasteless; it is heavier than air. What are its other physical properties? It is only sparingly soluble, water dissolving only 3% of it. Oxygen may be liquefied at −118° C. by a pressure of fifty atmospheres. The liquid has a pale steel-blue color, and boils at −181° C. under ordinary pressure.

37. Chemical Properties. — Oxygen combines with every known substance except fluorin, and is characterized by great chemical activity. It is the great supporter of com-

bustion. If both the oxygen and a combustible substance be *absolutely dry*, it has been shown that they will not combine. No satisfactory explanation of this fact has been offered. Oxygen is the only element capable of supporting respiration. Fish breathe the dissolved oxygen in water.

38. Uses. — Oxygen is necessary to animal respiration, to ordinary combustion, fermentation, and decay. It is used in the arts to increase the intensity of combustion, and is also used in medicine.

39. Burning in Air.

Experiment XXXI. — 1. Pour 10 cc. of lime water into a bottle containing air, shake the bottle, note the effect on the lime water; now, using a short piece of wire as a handle, lower a burning match into the bottle; when it has gone out cover with the hand and shake the bottle; note the changed appearance of the lime water. A milky appearance proves the presence of carbon dioxid.

2. Repeat the experiment using a bottle of oxygen.

When sulfur burns in air a gas having the characteristic odor of burning matches is formed.

3. Determine whether the gas formed when sulfur is burned in oxygen is the same that is formed when it burns in air, by burning sulfur in a bottle containing air and in one containing oxygen, and compare the odors of the gases formed. Discuss the relation between combustion in air and in oxygen.

The difference in activity is due entirely to the fact that in air oxygen is diluted with another gas which does not support combustion.

REVIEW QUESTIONS

1. Describe the preparation of oxygen from potassium chlorate. Mention precautions to be observed.
2. What is the office of manganese dioxid in the above process?
3. What are the tests for oxygen?

4. Compare the action of oxygen on charcoal at ordinary temperatures with its action at high temperatures.

5. Compare the product obtained by burning charcoal in oxygen with the product obtained by burning it in air.

6. Compare the action of oxygen on phosphorus at ordinary temperatures with its action at high temperatures.

7. What can you say of the products of combustion in air and in oxygen?

8. Discuss the occurrence of oxygen in nature.

9. State the physical properties of oxygen; the chemical properties.

10. Does oxygen occur uncombined in nature?

11. Mention several compounds containing oxygen which occur in nature.

12. Does oxygen display greater energy at high temperatures than at low temperatures?

CHAPTER V

COMBUSTION

40. Ordinary Combustion. — In its broadest sense, the term combustion is applied to all cases of chemical action which are accompanied by an evolution of heat and light. In the majority of cases, however, oxygen is one of the elements concerned in combustion, and because of the rarity of the exceptions, the term is sometimes defined as the union of a substance with oxygen, accompanied by the evolution of heat and light; and the classification of substances as combustible and incombustible depends upon this definition of the term. Thus a combustible substance is one which unites with oxygen with evolution of light and heat, and an incombustible substance is one which cannot unite with oxygen.

Many substances are products of combustion; thus water is composed of hydrogen and oxygen, and carbon dioxid of carbon and oxygen. In these compounds the hydrogen and the carbon have already combined with oxygen, and cannot directly combine with more.

41. Kindling Temperature. — A wise provision of nature makes it necessary to raise the temperature of substances slightly above that which ordinarily obtains, to cause them to combine rapidly with oxygen. If this were not true we should have no fuels. Substances differ widely in the temperature to which they must be raised to cause them to combine with oxygen, but for each combustible substance

there is a definite temperature at which it combines with oxygen with sufficient energy to develop heat and light, and this is called the kindling temperature.

If the kindling temperature of a substance is below the ordinary temperature, it will take fire when it comes in contact with the air, and must, therefore, be kept out of contact with the air. Such substances are said to be spontaneously inflammable. Several substances have kindling temperatures below a red heat, *e.g.* the gaseous hydrogen phosphid may be ignited with a test tube containing boiling water, and the vapor of carbon disulfid may be ignited with a glass rod heated to 120°. Most solid fuels require a temperature slightly above redness, while the diamond must be raised to nearly a white heat before combustion begins. In starting a fire we take advantage of differences in the kindling temperatures of substances. For example, paper is easily ignited, but the heat which it develops cannot ignite the anthracite; hence we often put charcoal between the paper and the coal, as paper can ignite the charcoal. The use of a coating of sulfur or paraffin on matches, to enable the phosphorus to ignite the wood, is another instance of the use of a substance having an intermediate kindling temperature.

The temperature produced by the combustion of a substance is not necessarily the same as its kindling temperature. In all cases of ordinary combustion the temperature produced is higher than the kindling temperature of the substance; burning particles thus raise adjoining particles to the kindling temperature, and the burning continues without further application of heat when once started.

There are, however, numbers of cases in which the temperature developed by the combustion cannot proceed without the continuous application of heat. The heat of the

electric spark ignites nitrogen, but the heat developed does not kindle the adjacent particles.

The facility with which a combustible substance may be ignited depends upon the quantity of heat, *i.e.* upon the number of heat units required to raise it to its kindling temperature. But, as we learn in physics, the temperature to which a substance is to be raised is only one of four quantities which determine the number of heat units required; the other three being the specific heat of the substances, its mass, and the number of heat units lost by conduction and radiation.

The amount of carbon to be kindled in a given stove depends upon the specific gravity and the porosity of the fuel; for example, charcoal, gas coke and anthracite coal are each of them nearly pure carbon, but they require very different amounts of kindling to ignite them. The specific gravity of the solid portions of these fuels are as follows:—

Pine charcoal	.40
Gas coke	.86
Anthracite	1.60

while the cell space or porosity expressed in cubic centimetres in 100 grammes of the fuel is as follows:—

Pine charcoal	200.4
Gas coke	60.
Anthracite	3.6

We thus see why charcoal requires comparatively little kindling to ignite it, although its kindling temperature is the same as the others.

The amount of heat lost by conduction has an important bearing on the amount of kindling required to build a fire. If the fuel is a good conductor of heat, it will be diffused throughout the mass, and such fuels are more readily

ignited if they are in small pieces, *e.g.* shavings are easier to ignite than a block of the same kind of wood.

Experiment XXXII. — Hold the laboratory burner horizontally over a sheet of white paper. Sprinkle some fine iron filings through the flame. What occurs? Pick up a few of the larger pieces on the paper and drop them through the flame again. What particles are raised to incandescence? Why?

Masses of metal in contact with the fuel often occasion considerable loss of heat by conduction. This action is nicely illustrated in the following experiment.

Experiment XXXIII. — Light a candle, bring a piece of wire gauze slowly down on the flame until it touches the wick. What occurs? Note the conditions above and below the gauze. Hold a lighted match above the gauze. What occurs? Explain. To what extent is the gauze heated? The Davy safety lamp used by miners illustrates this action.

42. Heat of Combustion. — It must be clearly understood that the light produced by combustion is due to the fact that the chemical action develops heat more rapidly than it can escape, thus raising the body to incandescence. There are many cases of oxidation, however, which take place slowly, or in which the substance is so situated that the heat is conducted away as fast as developed and in which a high temperature is not reached. The most important illustration of this action is the oxidation occurring within our bodies, which supplies the heat necessary to our existence. Other illustrations are found in the heat developed in compost heaps, in hotbeds, in the decay of wood, in cases of "spontaneous combustion," and in the rusting of iron.

The higher temperature acquired by a substance when it burns is readily accounted for by the difference in the *rate* at which it combines with oxygen. When two substances, such as carbon and oxygen, combine, their chemical affinity causes the atoms to rush toward each other, and the col-

lision which ensues increases their rate of vibration; that is to say, it develops heat. The *amount* of heat developed by the collision which forms a single molecule depends upon the magnitude of the attractive force and not upon the rate at which similar molecules are formed; it follows, therefore, that the oxidation of a given mass of a substance will develop exactly the same amount of heat when it burns that would have been developed if it had been oxidized slowly.

43. Chemical Energy. — All cases of direct chemical combination are due to attractions between unlike atoms; and whether the attraction be great or small, the collison of the atoms will develop heat. When molecules consisting of more than one atom act upon each other, the force which holds the atoms together in the original molecules must be overcome before a chemical change can occur; and the amount of heat developed in any reaction will accordingly depend upon the magnitude of the attractions of the atoms of the factors, as compared with the value of the attractions of atoms of the products. If the latter exceed the former, a chemical change will occur, accompanied by an evolution of heat. Chemical changes which evolve heat are known as *exothermic changes*, and those compounds which are formed from their elements by such changes are known as *exothermic substances.* Such substances are very stable, and when they are separated into their original elements the same quantity of heat that was evolved when they were formed disappears, or more exactly, is transformed into chemical potential energy. The formation of a much smaller class of substances is accompanied by the disappearance of heat; these are known as *endothermic bodies,* and when they are decomposed heat is evolved. They are usually unstable and often very explosive. All

endothermic substances possess chemical energy and can do work; that is to say, a substance which can combine with other substances without the aid of external energy possesses chemical energy. Much of the mechanical energy of the world is derived from endothermic substances, *e.g.* the fuels. The decomposition of carbon dioxid in the plant is an endothermic reaction in which the energy of the sunlight disappears. The carbon thus formed is stored up and may be again oxidized. For this reason the energy derived from wood and coal is sometimes spoken of as "stored sunlight."

44. Nomenclature of the Oxids. — The simplest chemical compounds are those composed of two elements only; they are known as *binary* compounds. Many binary compounds end with the letters *id;* but this rule cannot be depended upon in all instances.

Binary compounds of oxygen are called oxids; they are very numerous; *e.g.* oxygen forms five distinct compounds with nitrogen.

When there are two oxids of the same element it is quite common to distinguish them by adding the suffix *ic* to the name of the element to denote the oxid having the greater amount of oxygen, and the suffix *ous* to the name of the element to denote the oxid having the smaller proportion of oxygen. Thus, mercuric oxid has a larger percentage of oxygen than mercurous oxid, and nitric oxid a larger percentage than nitrous oxid.

If there are more than two oxids of the same element, prefixes are often used. Thus a *peroxid* contains a larger percentage of oxygen than the oxid to which the suffix *ic* is applied. Nitrogen peroxid, which contains a larger proportion of oxygen than nitric oxid, illustrates this usage.

A more scientific and simpler method of naming oxids

has been suggested, and is quite generally used. According to this plan the first part of the name of the oxid consists of the name of the element oxidized, and the second part of the name indicates the number of atoms of oxygen which the oxid contains, by the use of certain prefixes derived from the Greek. Oxids containing one atom of oxygen are called *monoxids, e.g.* carbon monoxid; those containing two atoms *dioxids, e.g.* carbon dioxid; those containing three atoms, *trioxids, e.g.* sulfur trioxid, etc.

45. Oxids in Nature. — Water, or hydric oxid is the most abundant oxid in nature, and sand (silicon dioxid) is next.

The ores of some of the most important metals are oxids, *e.g.* the red iron ore so common in this country is a compound of iron and oxygen, the molecule of which contains two atoms of iron and three of oxygen; and black iron ore, or lodestone, contains three atoms of iron and four of oxygen in its molecule. Many other ores are oxids, *e.g.* those of tin, manganese, etc.

REVIEW QUESTIONS

1. How do substances formed by burning in air compare with those formed by burning in oxygen?
2. Why is not combustion as rapid in air as in oxygen?
3. Define combustion. What are combustible substances?
4. Define kindling temperature. Which has the highest kindling temperature, sulfur, carbon, or phosphorus?
5. Mention examples of slow oxidation. How does slow oxidation differ from combustion?
6. Compare the amount of heat given off during slow oxidation and combustion.
7. What is meant by chemical energy? What substances possess it? What substances do not possess it?
8. From what source is the mechanical energy of wood derived? Explain.
9. What are oxids and how are they named? What do the terminations "ic" and "ous" indicate?

10. What important oxids occur in nature? Why are they so abundant?

11. Give evidences that a part of the air combines with the fuel in combustion.

12. Describe an experiment to show the relation of the weight of the products of a burning candle to the weight of the portion of the candle consumed.

13. How is combustion related to or distinguished from chemical action in general?

14. Mention conditions that favor combustion and chemical action in general.

15. Mention a condition favoring some chemical action but not combustion.

16. What is meant by kindling temperature? Explain the theory of shaving wood for use in starting a fire of the same kind of wood.

17. How does the chemical energy of the combustion of hydrogen compare with that of the combustion of other elements? Why?

18. Why is a fire of seasoned wood hotter than a fire of green wood?

19. Explain the use of sulfur in making the common friction match.

20. Why is a wood fire easily started with wood shavings?

21. Upon what does the temperature reached by combustion of a given quantity of fuel depend?

22. Mention five oxids occurring abundantly in nature.

23. Mention several substances which are acted upon by oxygen at ordinary temperatures.

24. Explain the effect of fine wire gauze when lowered over the flame of a lamp. Mention an important practical application of the principle involved.

25. Compare the kindling temperature of hydrogen with that of carbon. What bearing has their relative kindling temperature on the production of light by illuminating gas?

26. Explain the phenomenon of spontaneous combustion.

27. What would occur if the temperature developed by the combustion of nitrogen were higher than its kindling temperature?

CHAPTER VI

NITROGEN

SYMBOL N. — ATOMIC WEIGHT 14

46. Occurrence. — Nitrogen forms $\frac{4}{5}$ of the bulk of the air. It is found in combination in a large number of substances, *e.g.* in saltpetre or potassium nitrate, KNO_3, and Chili saltpetre, $NaNO_3$. It also occurs abundantly in ammonia, nitric acid, flesh, and other animal substances. Its compounds give to burned hair and woollens their peculiar odor. Many vegetable substances contain nitrogen, as cabbage, mushroom, horse-radish, and it is an essential constituent of quinine, morphine, prussic acid, and strychnine. It forms a part of nearly all explosives, as nitroglycerin, gunpowder, etc.

47. Preparation. — Nitrogen may be prepared by removing the oxygen from the air. Any method which burns up the oxygen of the air and forms solid or liquid products, yields nitrogen which is reasonably pure. If any of the products are gaseous they will be mixed with the nitrogen, which will therefore be impure.

First Method. — Introduce a jet of burning hydrogen into a bottle of air. After the flame is extinguished there will remain in the bottle, nitrogen and the product of the combustion of hydrogen — H_2O.

Second Method. — Phosphorus burns in the air forming phosphorus pentoxid, P_2O_5, a flaky white substance which is soluble in water.

Experiment XXXIV. (See Experiment 28). — Support a piece of chalk over water in the water pan by means of a wire standard. Make a hollow in the chalk, place a piece of dry phosphorus about the size of a pea in it. Ignite the phosphorus, quickly cover it with the large bottle so that the mouth of the bottle is under water. What chemical change takes place? Notice any change in the volume of the air. Explain. Does all the air support combustion? Take the bottle from the water pan (do not allow the water to escape) and shake it. What is the result? Does the white cloud which at first filled the bottle remain? Test the gas with a lighted taper. Is it combustible? Is it poisonous? State its physical properties.

Third Method. — If air be passed through a tube containing heated copper filings, the oxygen combines with the copper, forming copper oxide, and nitrogen may be collected. Nitrogen prepared from the air will contain the impurities which exist in the air. Pure nitrogen may be prepared as follows: —

Fourth Method. — Heat ammonium nitrite and collect evolved gas over water.

$$NH_4NO_2 = 2\,H_2O + 2\,N.$$

On account of the unstable character of ammonium nitrite, it is difficult to keep a supply on hand; in practice, therefore, a mixture of ammonium chlorid and sodium nitrite is usually substituted for the ammonium nitrite. When this mixture is heated the reaction proceeds according to the following equation: —

$$NH_4Cl + NaNO_2 = NaCl + 2\,H_2O + N_2.$$

This method supplies the purest nitrogen.

48. Physical Properties. — The last experiment taught us certain physical properties of nitrogen; mention them. The following additional physical properties not easily shown experimentally are worthy of consideration. It is sparingly soluble in water, only 1.6 % being dissolved at 10° C. It may be liquefied at − 193° under pressure of one atmosphere. It is slightly lighter than air.

49. Chemical Properties. — Nitrogen combines directly with very few elements, and combination with these elements is effected with difficulty. By indirect methods it can be made to combine with hydrogen, and with hydrogen and oxygen. Its chemical affinities are exceedingly feeble, and the compounds which it forms are very unstable. Gunpowder, nitroglycerin, and many other explosives are nitrogen compounds, and owe their characteristic properties to the ease with which they are decomposed. The rapid decay of animal and vegetable substances which contain nitrogen is a further illustration of the unstable character of nitrogen compounds.

REVIEW QUESTIONS

1. State the physical properties of nitrogen.
2. State its chemical properties, activity, combustibility, relation to explosives, relation to decay. Is it poisonous?
3. What proportion of the air is nitrogen? How is this shown?
4. Describe an experiment for obtaining nitrogen by the use of phosphorus. Give the name and formula of the fumes formed, and account for their disappearance.
5. Compare oxygen with nitrogen with respect to (*a*) chemical activity, (*b*) occurrence, (*c*) number of compounds, (*d*) relation to combustion and life, and (*e*) physical properties.
6. Why is nitrogen an important constituent of most explosives?
7. What are nitrogeneous foods?

CHAPTER VII

HYDROGEN

Symbol H. — Atomic Weight 1

50. Occurrence. — Hydrogen is never found uncombined in nature; its compounds, however, are widely distributed. It forms $\frac{1}{9}$ of the weight of water, and occurs in all animal and vegetable matter. It is the only substance common to all acids.

51. Preparation by the Action of an Acid on a Metal.

Experiment XXXV. — 1. Put a few pieces of granulated zinc in the test bottle. Cover them with dilute hydrochloric acid. What occurs ?

2. After a minute or two hold a lighted match over the bottle. What occurs ?

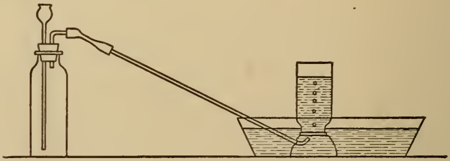

Fig. 4.

3. Put a few pieces of zinc in your generating bottle. In one hole in the rubber stopper put a straight glass tube long enough to reach to the bottom of the bottle ; in the other fit a bent tube with a delivery tube attached. Pour enough dilute sulfuric acid into the bottle to cover the zinc. Collect the gas over water.

Caution. — The gas is explosive when mixed with air; when pure, it burns quietly. To determine when all the air which filled the bottle at the beginning of the experiment is driven off, collect a small bottle of gas; when full, raise it from the water, mouth downward, and apply a match. If the first bottle of gas explodes, repeat until a bottle of gas is obtained which burns quietly.

4. Collect three bottles of pure gas.

5. Place the first bottle, mouth upward and uncovered, on the table. After a few minutes, test it, to see whether or not it contains hydrogen. Is hydrogen lighter or heavier than air?

6. Pour the hydrogen in the second bottle upward into an inverted bottle containing air. Test each bottle with a match. Is there any hydrogen in the inverted bottle? In the other bottle?

7. Light a candle with a wire attached for a handle. Hold the third bottle mouth downward and thrust the lighted candle well into the bottle. What occurs? What burns? Does the candle burn? Withdraw the candle slowly. Is it alight? Why? Put it back into the hydrogen. Does hydrogen support combustion? Is the mouth of the bottle heated?

Experiment XXXVI. The Philosopher's Lamp. (Optional.) — Remove the delivery tube and substitute for it a tube drawn out to a fine point. If you are sure that the gas is pure, *i.e.* if you have not taken the stopper out of the generating bottle since testing the gas, light the gas at the end of the pointed tube. Hold a cold dry bottle over the flame. What do you see in the bottle? Where did it come from? (Chemical examination proves it to be pure water). What is the product of the combustion of hydrogen?

Hydrogen may be prepared by several other processes; for example, by decomposing water by electricity (Experiment 40); by decomposing water by metals at ordinary temperatures (Experiment 42); by passing steam over heated metals (Art. 72, etc.).

The following equations represent chemical changes in the preparation of hydrogen by the action of an acid on a metal: —

$$Zn + 2\,HCl = ZnCl_2 + 2\,H.$$

This is the reaction when hydrochloric acid is used. If sulfuric acid is used, the following equation expresses the reaction: —

$$Zn + H_2SO_4 = ZnSO_4 + 2\ H.$$

52. Physical Properties. — Pure hydrogen is odorless, and is the lightest known substance; one litre of it at ordinary pressure weighing .08956 gramme. It may be liquefied by extreme cold and pressure, but is more difficult to liquefy than any other gas. It diffuses more rapidly than any other gas. Water dissolves only 2% of hydrogen.

53. Chemical Properties. — In its chemical affinities hydrogen closely resembles a metal; it has a strong affinity for oxygen, chlorin, and a few other elements, and the compounds which it forms with carbon indirectly are very numerous; it is, however, very difficult to get it to combine directly with carbon.

54. Comparison of Physical and Chemical Properties of Hydrogen and Oxygen. — Hydrogen will burn; oxygen supports combustion. Hydrogen has affinity for few substances; oxygen for many. Hydrogen is the lightest known substance. Oxygen combines readily with carbon, sulfur, phosphorus, and iron. It is difficult to get any of these elements to combine with hydrogen. The two elements have opposite chemical properties; yet in their physical properties they resemble each other.

55. Uses. — On account of its great affinity for oxygen, hydrogen is extensively used for the purpose of extracting oxygen from compounds containing it, *i.e.* as a reducing agent.

56. Heat and Chemical Energy of the Combustion of Hydrogen. — The chemical affinity of hydrogen for oxygen is

greater than that of any other known substance. The heat produced by the combustion of hydrogen is therefore greater than that of any other substance. One pound of hydrogen in burning gives off 34,400 heat units; that is, it develops enough heat to raise 34,400 pounds of water from 0° C. to 1° C.

The oxyhydrogen blowpipe consists of a tube *H* (Fig. 5), through which hydrogen flows, and at the end of which it is ignited. In the centre of this is a smaller tube

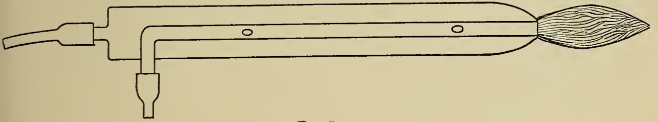

Fig. 5.

through which a stream of oxygen is forced into the flame. The flame produced gives very little light, but its temperature is between 2000 and 2200° C.; it is, therefore, used in working platinum and other metals fused with difficulty. A piece of lime held in the flame is heated to incandescence, and emits a bright light equivalent to about 120 standard candles. This device is known as the calcium light.

Experiment XXXVII. — Take notes on the effect of the oxyhydrogen blowpipe flame upon bits of lead, zinc, copper, steel, iron, glass, and calcium oxid.

57. Burning of Oxygen or Air in Hydrogen. — If a jet of oxygen or air be introduced into a vessel containing hydrogen, the oxygen or air may be ignited and will burn as readily as hydrogen burns in oxygen or in air.

If a stream of hydrogen be passed through the tube *H* (Fig. 6), and ignited at the bottom of the bottle, a jet of

air introduced through the hydrogen flame will burn with a flame of the same character as that produced when hydrogen burns in air.

58. Product of this Combustion. — The combustion of hydrogen must form an oxid of hydrogen. Water is an oxid of hydrogen, and analysis of the moisture condensed on any cold object held over the hydrogen flame, proves that water is the product.

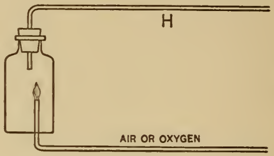

FIG. 6.

59. Formation of this Substance in Ordinary Combustion. — Nearly all fuels contain hydrogen, and therefore form more or less water when they burn; this can be shown by holding a cold object over the flame. Moisture can thus be condensed over burning oil, wood, coal, gas, etc. When oxygen combines with the waste products of the body in the lungs, the hydrogen of the products is turned into water; thus the well-known cloud formed by the breath in cold weather is this moisture rendered visible by condensation. This may be easily shown by breathing upon any cold dry object.

REVIEW QUESTIONS

1. Explain the cause of the moisture which appears on a lamp chimney when the lamp is lighted.

2. Why does this film disappear?

3. If water is a product of combustion, why does it not extinguish the fire?

4. State the symbol, atomic weight, and occurrence of hydrogen.

5. Describe the process of preparing hydrogen from zinc and hydrochloric acid. Write the reaction.

6. Discuss the physical and chemical properties of hydrogen.

7. Describe the oxyhydrogen blowpipe. For what is it used?

8. Show how a jet of air may be burned in hydrogen.

9. What does the moisture which gathers on a cold object held over a lighted kerosene lamp indicate as to the composition of the kerosene oil?

10. Discuss the heat and chemical energy of the combustion of hydrogen.

11. State the color and odor of the gas prepared in Experiment 35.

12. Is hydrogen explosive? Under what conditions?

13. Does hydrogen combine with oxygen at ordinary temperatures? at high temperatures?

14. Describe the hydrogen flame as to (*a*) color, (*b*) amount of heat, (*c*) amount of light.

CHAPTER VIII

CHEMISTRY OF WATER

FORMULA H_2O. — MOLECULAR WEIGHT 18

60. Occurrence of Water in Nature. — Three-fourths of the earth's surface is covered with water. It exists in the atmosphere, in all vegetable and animal matter, in the soil, and even in the rocks.

Experiment XXXVIII. — Heat a small piece of alum in a test tube. Note evidence that it contains water. Is the loss of water accompanied by a change in the crystal? Explain. Repeat the experiment, using pieces of gypsum, meat, potato, etc.

61. Properties of Water. — At ordinary temperatures pure water is a tasteless, odorless, transparent fluid, colorless in thin layers, but distinctly blue when viewed in large masses. At its greatest density water is 773 times heavier than air.

Many of the properties of water are used as standards by means of which we may express the corresponding properties of other substances. The specific gravity of all solids and liquids expresses the relation between the weight of the substances and the weight of a like volume of water.

The specific heat of all substances is similarly based upon that of water. In the metric system the unit of weight is the weight of a cubic centimetre of water, and the melting and boiling points of water are the standard temperatures used in the manufacture of thermometers.

62. Solution. — Water dissolves a greater number of solid, liquid, and aeriform substances than any other solvent.

The adhesion between the molecules of the dissolved substance and those of water overcomes the cohesion of the substance dissolved, and it diffuses through the mass of water, forming a transparent *solution* in which the dissolved substance is invisible. Strongly colored substances, however, impart their characteristic color to the solution. There is a limit to the amount of a substance which a solvent can dissolve at a given temperature and pressure, and a solution which contains all of a substance which it can dissolve is said to be *saturated*.

We can now understand why solution aids chemical action. The molecules are no longer held firmly together by cohesion; they are free to move, and are thus easily brought into the intimate contact necessary to chemical action by their chemical affinities.

If sugar is dissolved in water the solution seems to be simply a mechanical mixture; there is no evidence that a chemical change has taken place; and if the solution is evaporated the sugar is recovered unchanged. This, therefore, is a *physical solution*.

In Experiment 9, marble was dissolved in hydrochloric acid and a chemical change was proven. Such solutions are called *chemical solutions*.

63. The Effect of Heat on the Solution of Solids and Gases.— Physics teaches us that whenever a solid changes to the liquid form a certain quantity of heat is rendered latent. Hot water can supply this heat more readily than cold, and therefore solids are more rapidly dissolved by hot than by cold water. Furthermore, a larger quantity of most solids is dissolved by hot water than by cold. At high temperatures the cohesion of the molecules of a solid is less than at low temperatures, therefore there is less force to overcome

in dissolving it. As the water cools, latent heat is withdrawn from a certain quantity of the dissolved substance, causing it to assume a solid state.

When a gas is dissolved in a liquid its latent heat must be absorbed, but a cold liquid can absorb heat more rapidly than a warm one, and therefore a gas is more rapidly dissolved in a cold liquid than in a warm one. As the temperature of a solution of a gas is raised, a portion of the heat is used to change the state of the dissolved gas and a portion is liberated in gaseous forms.

64. Water of Crystallization. — Experiment 38 taught us that crystalline alum contained water, and that the alum lost its geometrical form when the water was driven off. Many other crystals are like alum in this respect, and there is evidence that the water which they contain is held in feeble chemical combination. The water of crystallization does not make the substance moist, as it would if absorbed mechanically, and further, a given substance requires a definite amount of water for each molecule of the crystal. Certain substances form two or more kinds of crystal, requiring different quantities of water; and in some crystals the color depends upon the amount of water of crystallization. Cobalt chlorid is often used as a *sympathetic ink* because of the change in color produced by expelling the water of crystallization.

In some crystals the water is held so feebly that they lose either the whole or a portion of their water of crystallization when exposed to the air, and in so doing lose their particular geometrical form. This process is known as *efflorescence.*

Other crystals absorb water from the air and assume other geometrical forms, in some cases absorbing enough water to dissolve the crystal. Such crystals are said to be *deliquescent.*

Experiment XXXIX. — Put a crystal of ferrous sulfate and a small piece of calcium chlorid on separate pieces of paper and expose them to the air for several days. Which is efflorescent? Which deliquescent?

65. Hydroxids. — Strictly speaking hydroxids are compounds formed by replacing one atom of hydrogen in the molecule of water with an atom of another element or with a group of elements.

$$Na + H_2O = NaOH + H$$
$$CaO + H_2O = CaO_2H_2.$$

According to this definition most acids are hydroxids but chemists rarely apply the term to them, whereas, all chemists agree in calling a compound formed by the union of a metal with hydrogen and oxygen an hydroxid.

These compounds, which are very important and which are discussed more fully in Chapter XII., are sometimes called hydrates, but the term is rather objectionable because the termination *ate* is used to distinguish a class of compounds to which the hydroxids do not belong.

66. Electrolysis of Water.

Experiment XL. (Performed by the instructor.) — Take notes upon this experiment, answering all the following questions and describing the apparatus used. An electric current is passed through acidulated water from one lead or platinum electrode to another. Gas is evolved which is collected in two tubes. Where is the gas liberated? How does the volume over the positive electrode compare with that over the negative? What gas is collected over the positive electrode? How do you know? What gas is collected over the negative electrode? Does this experiment prove that water is composed of two elements and no more? Is the volume of the water decomposed equal to the volume of the gases formed? How do you know?

67. Synthesis of Water.

Experiment XLI. (Performed by the instructor.) — Eudiometer tube — a "U" shaped tube of glass about 18 inches long closed at one end, having two platinum wires inserted at opposite sides near the closed

end. Fill the eudiometer tube with mercury and invert in a mercuric bath. Introduce a certain amount of oxygen, removing the tube and bringing the mercury to a level in both arms. Read the amount of oxygen in the tube, filling the arm with mercury again. Introduce about twice as much hydrogen in a similar manner, determining the exact volume, placing the thumb over the end of the tube that is open, and wrapping the tube in a towel, pass an electric shock through the wire. An explosion occurs, and water is formed. Some of the gas remains in the tube. Testing this residual gas to determine whether it is hydrogen or oxygen and subtracting its volume from the quantity used, we determine the volume of the two gases which combined.

This experiment proves that there are only two elements in water.

68. Formation of Water by passing Hydrogen over Heated Oxid. — When mercuric oxid is heated, oxygen is liberated. If a stream of hydrogen be passed over a heated oxid, the hydrogen and oxygen combine to form water. When copper oxid is used, the following reaction takes place: —

$$CuO + 2\,H = Cu + H_2O.$$

In this experiment if the weight of water formed be determined and the tube containing copper oxid be weighed before and after the heating, it will be found that $\frac{8}{9}$ of the weight of water came from the copper oxid, thus proving that $\frac{8}{9}$ of the weight of water is oxygen.

69. Composition of Water by Weight and by Volume. — Preceding experiments have shown that water contains twice as much hydrogen as oxygen, *volume* alone considered, and that it contains eight times as much oxygen as hydrogen, *weight* alone being considered. These two seemingly contradictory facts being proven, it follows that the single volume of oxygen must be eight times as heavy as the two volumes of hydrogen, and that equal volumes being considered, the oxygen is sixteen times as heavy as the hydrogen. In considering the composition of a compound, care must be taken to distinguish between these two

methods of stating the composition, and it must be remembered that the volumes used are, in all cases, the volume in an aeriform state, and not the solid or liquid state. One further fact should be stated here: when two volumes of hydrogen combine with one volume of oxygen to form water they do not form three volumes of steam; their volumes is condensed one-third, so that —

2 vols. of hydrogen + 1 vol. of oxygen form 2 vols. of steam,

and this is the way in which the composition of a substance by volume should be stated. For further discussion of this topic see p. 118.

The composition of a substance by weight is the same in the solid as it is in the liquid and aeriform state. That of water may be stated thus: —

2 parts (by weight) of hydrogen + 16 parts of oxygen form 18 parts of water.

70. Decomposition of Water by Metals.

Experiment XLII. — 1. Fill a medium sized bottle with water and invert it in the yellow dish, which should be about half full of water.

2. Thoroughly dry a small piece of wire gauze in the gas flame, roll around a lead pencil, so that it forms a cylinder that is double walled in all parts. Fold one end of the cylinder over and pinch the bend with a pair of pliers. Drop a piece of sodium the size of a pea into it, using the pliers as before.

3. Lift the bottle slightly, but not enough to allow the water to escape, and thrust the wire gauze beneath it. After all action has ceased slip a glass plate under the mouth of the bottle and remove the bottle from the water pan, placing it right side up on the table.

4. Test the gas with a burning match. Do you recognize the gas? How? Where did it come from? Drop a piece of pink litmus paper in the bottle. Does it change color? Does water from the laboratory faucet produce a similar change? What became of the sodium.

Explain concisely all that has occurred. The reaction is as follows: —

Sodium + Water = Sodium Hydroxid + Hydrogen

$$Na + H_2O = NaOH + H$$

NOTE. — Litmus paper is prepared by dipping strips of paper into an infusion of litmus. It turns red, when treated with an acid, and blue, when treated with an alkali.

72. Decomposition of Water by passing Steam over Heated Metals. — Certain other metals which decompose water slowly or not at all at ordinary temperatures, decompose it easily at high temperatures. If steam be passed through a tube containing bits of iron heated to redness, it will be decomposed and hydrogen may be collected over water.

$$3\,Fe + 4\,H_2O = Fe_3O_4 + 8\,H.$$

QUERY. — How may the weights of hydrogen and oxygen resulting from the decomposition be determined in this experiment?

73. Water Gas. — At high temperatures carbon also decomposes water, and this fact is the basis of the process of manufacturing water gas. Steam is passed over highly heated coal or coke (carbon), which combines with the oxygen of the water, forming carbon monoxid (CO). The reaction is as follows: —

$$C + H_2O = CO + 2\,H.$$

Both carbon monoxid and hydrogen are combustible gases, they are both odorless, and burn with feebly luminous flames. To overcome these objections it is customary to enrich water gas by adding certain gases, obtained by decomposing naphtha, which give the gas a distinct odor, and which greatly increase the amount of light produced. The illuminating gas used in many of our cities is prepared by this process, but the increased value of the ammonia and of the tar obtained as by-products in the process of manufacturing illuminating gas from bituminous coal has rendered the economy of water gas questionable. As water gas is more poisonous than ordinary illuminating gas, laws have been passed in certain states prohibiting its use.

74. Oxidation and Reduction. — The process of abstracting oxygen from a body is called *reduction*. In the manufacture of water gas the hot carbon abstracts the oxygen from the water, illustrating its use as a reducing agent. Carbon is extensively used in the reduction of ores to a metallic state. The union of a substance with oxygen is called *oxidation*, and the reagent which causes the oxidation is called the *oxidizing agent*. Nitric acid and potassium chlorate are excellent oxidizing agents, as will be noticed in several subsequent experiments. In Experiment 42 the sodium was oxidized by the water, but water is not among the better oxidizing agents.

75. Natural Waters. — Absolutely pure water is never found in nature. The impurities which it contains are of two classes: first, the *inorganic*, or those derived from the rocks; and second, the *organic*, or those derived from the decay of animal matter or vegetable substances. Some of the impurities are held in solution, while others are suspended and carried along by moving water. The purest water found in nature is rain water, particularly that which falls in country districts after it has been raining some time. But even rain water contains impurities; as it falls through the air it washes it, removing those suspended matters which are always present in it, and dissolving small quantities of the gases of the atmosphere. As soon as the rain reaches the earth, its great solvent power is exerted upon the mineral matter with which it comes in contact, and it becomes more impure. An impure water is not necessarily unfit for household purposes. See Article 78.

76. Spring Water always contains dissolved mineral matter as well as a considerable quantity of carbon dioxid derived from the decomposition of plants. Waters flowing

through different strata would naturally contain different amounts and different kinds of minerals, and we therefore have various kinds of spring waters. *Sulfur springs* usually issue from rocks containing a decomposing sulfid. A line of sulfur springs across western New York marks the outcrop of the Hamilton shale which, in certain layers, contains a great deal of iron sulfid. The *chalybeate springs* contain some compounds of iron in the same way, and the *effervescent waters* have some gas in solution. The famous springs at Saratoga, N.Y., belonging to this class, contain carbon dioxid.

River water differs from spring water because a part of it, at least, has not been filtered through porous rock and thus relieved of suspended matter. It is consequently turbid, while spring water is usually clear and sparkling. Sea-water contains a large amount of mineral matter, the average quantity being about 3.5 % of its weight.

77. Hard Water. — Certain salts which are frequently present in natural waters prevent the formation of a lather with soap in the ordinary process of washing, and give to waters containing them the property known as hardness. The chief substances which produce this effect are the compounds of calcium and magnesium. Soap is decomposed by such waters, forming an insoluble, curdy precipitate or scum which prevents the cleansing action of the soap until all of the hardening salts have been removed. Hardness due to the presence of carbonates may be removed by boiling the water, and is called *temporary hardness.* Waters having this kind of hardness are common in limestone regions. The limestone is an impure calcium carbonate ($CaCO_3$), and is not soluble in water, but water containing carbon dioxid converts the calcium carbonate into an acid carbonate ($CaH_2(CO_3)_2$) which is soluble. *Permanent hard-*

ness, or that which remains after prolonged boiling, is usually due to the presence of sulfates.

78. Potable Waters. — The inorganic impurities found in natural waters are very rarely injurious to health. The organic impurities, however, are usually accompanied by living germs, or bacilli, by means of which such diseases as cholera, typhoid fever, diphtheria, etc., are propagated. The principal source from which these dangerous organic impurities are derived is the drainage of houses and villages; and though waters thus contaminated may become pure again through the action of the air and sunlight, it is not safe to rely upon this method of purification in the water which is to be used for drinking purposes.

It is of the greatest importance that the water used for drinking purposes should be as pure as possible; to this end it should be frequently tested, and if there is the slightest suspicion of contamination, it should be thoroughly boiled.

In general it may be assumed that springs, deep wells, and mountain rivers and lakes, are safe sources of water supply; that stored rain water and surface water from cultivated land are unreliable; and that shallow wells and rivers, to which sewage gains access, are dangerous sources.

79. Distillation and other Methods of Purification. — The best method of purifying water is by distillation. This process removes both organic and inorganic impurities; and, when properly conducted, supplies a perfectly pure water. On shipboard the salt water of the ocean is distilled.

Experiment XLIII. — Using a flask, distil 30 or 60 cc. of water. Under "Physical Properties" note color, taste, odor. What action has it upon litmus paper? Does it leave a residue upon evaporation? Answer same questions concerning some natural water.

Boiling. — Thoroughly boiling water, renders most of the organic impurities harmless. Disease germs are destroyed

and albuminous matter coagulated so that it may be readily removed by filtration.

Filtration does not remove the dissolved inorganic impurities, — these can be removed only by chemical processes (see next article); but when properly conducted it removes both organic and inorganic suspended matter, including disease germs.

Small *charcoal* or *sand filters* are more apt to contaminate the water passing through them than to purify it; for after a few days' use, the filter becomes so saturated with germs that the filtered water contains more of them than it did before it was filtered. Unless the filter is so constructed that the charcoal or sand may be removed and exposed to air and sunlight, it is unsafe to use it. The wire strainers, sometimes called filters and made to be attached to faucets, may remove some of the suspended matter, but they do not remove the germs. The *Pasteur filter*, in which the water passes through a natural stone, is the most efficient small filter; the stone must be removed and boiled, or exposed to the air for oxidation of the organic matter, occasionally, otherwise germs will pass through the filter after a time.

Chemical Methods. — Attempts are rarely made to purify water on a large scale by chemical processes. But such processes have been employed for many years by frontiersmen and in regions where no drinkable water exists.

If alum is gradually added to impure water containing calcium carbonate, calcium sulfate is formed, carbon dioxid is evolved, and aluminum hydroxid is precipitated. The aluminum hydroxid entangles the organic matter present and settles with it to the bottom, leaving the water clear and sparkling. If the water contains an insufficient amount of calcium carbonate, it will not be rendered perfectly clear, and the deficiency must be supplied by adding sodium carbonate.

Ferric chlorid, iron borings, and potassium permanganate have each been used successfully in chemical processes of purifying water.

80. Softening. — (*a*) *Temporary Hardness.*

Experiment XLIV. — 1. Dissolve 1 gramme of soap shavings in 10 cc. of distilled water in a test tube. Draw out one end of a glass tube, about $\frac{1}{4}$ inch in diameter, to a point thus: —

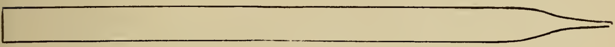

FIG. 7.

This is to be used as a dropping tube.

2. Test 10 cc. of hard water, made by passing carbon dioxid through lime water until it is clear, in a test tube as follows: Using the dropping tube, add a single drop of the soap solution to the hard water, shake the tube, repeat the operation as often as may be necessary to determine the number of drops required to produce frothing.

3. Test 10 cc. of distilled water in the same way. What do you conclude concerning the relative values of hard and pure water for washing purposes?

4. Boil some of the hard water and test 10 cc. as before. How does this compare in value with the unboiled sample?

Boiling decomposes the soluble carbonates, expelling carbon dioxid and precipitating the insoluble carbonate. The "fur" which forms on the bottom and sides of the tea-kettle and the scale which forms on the shell and tubes of a steam boiler are each due to the repeated removal of carbonates from water by boiling.

Clark's process of removing temporary hardness is quite extensively used by water-works engineers. The hardness is estimated, and the quantity of "milk of lime" required to transform the amount of soluble carbonate present into insoluble carbonate is then added. The reaction is as follows: —

$$CaH_2(CO_3)_2 + CaO = H_2O + 2\ CaCO_3.$$

(*b*) *Permanent Hardness.*

Experiment XLV. — 1. Test 10 cc. of hard water, made by dissolving calcium sulfate in distilled water and filtering, as in Experiment 44. Note the number of drops of soap solution necessary to produce frothing. Boil some of the above-mentioned hard water, and test as before. What effect does boiling have on permanent hardness?

The sulfates and chlorids of lime and magnesium are decomposed by sodium carbonate (common washing soda) with the following reaction: —

$$CaSO_4 + Na_2CO_3 = CaCO_3 + Na_2SO_4.$$

The sodium sulfate produced has no effect on the soap, but water containing it should not be used for drinking purposes. Hence the use of washing soda for softening water.

81. Natural Methods of Purification. (*a*) *Action of the Air.* — Disease germs die when exposed to the sunlight, and organic impurities are oxidized when exposed to the air. Therefore, impure water running through shallow streams, or over precipices in a thin sheet, tends to become pure again.

(*b*) *Filtration* through beds of sand or porous rock removes suspended matter. It is in this way that most spring waters are rendered clear and transparent; it should be remarked, however, that clearness is not unfailing evidence that water is healthful.

(*c*) *Sedimentation.* — In ponds and lakes water is often rendered clear, the suspended matter settling to the bottom under the influence of the force of gravity.

(*d*) *Natural Distillation.* — This is the most important method of purifying the water in nature. The sun is the source of heat, and since evaporation takes place at a temperature far below the boiling point, it occurs everywhere and at all times, even in winter, and the amount evaporated every hour is enormous. The vapor rises to the upper atmosphere, and encountering cold currents is

condensed in microscopic particles which float in the air. These increase in size as they move about, and finally fall as rain. The purest water found in nature is rain water, particularly after it has rained some time.

Mountain streams which flow over rocky beds, notably those which flow over beds of sandstone, have exceptionally pure waters. Water which flows over limestone dissolves some of the stone and becomes hard.

NOTE. — It is suggested that students take advantage of the opportunity offered by Experiments 46–48 to examine the water which they habitually use for drinking purposes.

82. Tests for Organic Impurities.

Experiment XLVI. — Fill a tall glass jar with water to be tested. Add a few drops of sulfuric acid, then add a weak solution of potassium permanganate, drop by drop, until the water assumes a violet tint. If organic matter be present, the color gradually grows lighter. If the color remains unchanged for an hour, the water may be considered safe for drinking purposes.

Nessler's Test. — Nessler's reagent is prepared by mixing potassium iodid and mercuric chlorid, and adding caustic soda. It furnishes a very delicate test for free ammonia, which is evidence of decomposing organic matter.

A drop or two of the reagent added to water containing ammonia gives it a brown color; the greater the amount of ammonia, the darker the shade of brown. In practice it is customary to concentrate the ammonia in 500 cc. of the water to be tested by distilling it with a small quantity of sodium carbonate, and testing the first 50 cc. of the distillate with Nessler's reagent.

Experiment XLVII. Test for Chlorids.

NOTE. — The members of the class should secure samples of drinking water from as many sources as possible, including water from a shallow well, a deep well, a pond, and a stream. Distilled water and a solution of common salt will be required; these are for comparison only, as distilled water contains no chlorin, and salt water a great deal. (Common salt is composed of sodium and chlorin.)

1. Add a few drops of nitric acid, free from chlorin, to 25 cc. of the distilled water, then add a few drops of silver nitrate solution.

2. Treat 25 cc. of the salt water in the same way. Compare this

with the distilled water. How can you distinguish water containing chlorin from pure water?

3. Concentrate 50 cc. of the drinking water to be tested to 25 cc. by boiling, and repeat the test.

Does it contain chlorin? Compare your result with that of other members of the class, and answer the following questions: —

Does the water from the shallow well become milky? Does that from the deep well? Which samples contain chlorin?

Sewage always contains chlorids, and hence if a drinking water is found to contain a chlorid, it is to be suspected, and the chlorin must be *proved* to come from some other source, or the water should be avoided. If a well is sunk near the sea, or near a deposit of rock salt, its water may be perfectly wholesome although containing a relatively large amount of chlorids.

83. Tests for Inorganic Impurities. — The presence of inorganic impurities is usually made known without a chemical test; the presence of hydrogen sulfid is detected by the odor, free carbon dioxid by effervescence, iron by the taste, and dissolved salts by the action of soap.

The presence of solids in solution may be determined by evaporating a few drops of the water to be tested on clean platinum foil. If solids be present, a residue will remain on the foil.

The fact that one usually desires to know concerning the inorganic impurities of water supplied for household use is its degree of hardness.

Experiment XLVIII. To determine the Degree of Hardness. — Pour 70 cc. of the water under examination into a flask. Add 1 cc. of "Clark's soap solution," insert a stopper, and shake thoroughly. Set it aside for two or three minutes; if a lather does not remain on the surface of the water at this time, add a second cubic centimetre of the soap solution. Repeat this process until a permanent lather is obtained. The number of cubic centimetres of soap solution used is equal to the number of degrees of hardness, and is one greater than the number of grains of calcium carbonate per imperial gallon.

84. Hydrogen Dioxid, H_2O_2. — This is a colorless bitter liquid somewhat heavier than water. It is very unstable, the molecule breaking up into water and oxygen. Because of this property it is manufactured on a large scale for use as a bleaching agent. It is also used as a disinfectant, and to some extent in laundries; it bleaches the skin and hair, and rapidly oxidizes metals.

REVIEW QUESTIONS

1. Describe the compounds which hydrogen forms with oxygen.

2. Describe two methods of decomposing water, and state how the weight of each gas may be determined in each case.

3. Mention three ways in which water may be decomposed, and two ways in which water may be formed.

4. Describe the electrolysis of water, and state how you would determine which gas is oxygen.

5. Describe an experiment illustrating the preparation of water by passing hydrogen over heated oxids.

6. What is an analytic experiment? — a synthetic experiment?

7. Describe experiments which prove (*a*) that water is composed of oxygen and hydrogen, (*b*) that it contains no other element, (*c*) that it contains twice as much hydrogen as oxygen by volume.

8. Describe the manufacture of water gas. Give equation; compare it with other illuminating gas.

9. Distinguish between an oxidizing agent and a reducing agent. Give examples.

10. Describe an experiment illustrating (*a*) oxidation (*b*) reduction.

11. What is meant by the term "hard water"? Explain how hard water may be made soft.

12. Distinguish between permanent and temporary hardness.

13. Give at least two tests by which impurities may be detected in water which appears pure to the eye.

14. How may water containing organic impurities be so purified as to render it safe to use?

15. Describe three ways in which water is purified by nature.

16. Describe the Pasteur water filter, and compare it with some other filter as to efficiency and as to necessary care.

17. Account for the formation of caves in limestone regions.

CHAPTER IX

PROBLEMS

85. Composition by Weight. — For the present purpose we will assume that the formulas of many substances, not yet examined by the pupil, are known. The formula of each compound shows the kind of atoms and the number of each kind in a molecule of the substance. Considering the atomic weight of each element, we may determine from the formula the number of parts, by weight, of each element in the compound. Thus, in the case of water, H_2O, the formula tells us that each molecule contains two atoms of hydrogen and one of oxygen; the two atoms of hydrogen weigh two microcriths, and the atom of oxygen weighs 16 microcriths; the molecule of water, therefore, weighs 18 microcriths, and $\frac{2}{18}$ of the weight of the molecule is hydrogen, and $\frac{16}{18}$ oxygen; further, since all molecules are exactly alike, $\frac{2}{18}$ of the weight of any number of molecules of water must be hydrogen, and $\frac{16}{18}$ oxygen.

The molecular weight of any substance is the sum of the weights of all the atoms in the molecule. Therefore, to find the molecular weight of a given molecule, add the number of microcriths of each element composing the molecule.

Case I

86. To compute the Weight of an Element in a Given Weight of a Compound. — To solve these problems it is merely necessary to know how many microcriths of that

element there are in the total number of microcriths in the molecule; since the ratio of the weights in the total amount of the substance is the same as in the molecule.

Example. — How much oxygen in 100 lbs. of sulfuric acid?

From the formula of sulfuric acid, H_2SO_4, we can (knowing the atomic weights of each of the elements) calculate that the molecule weighs 98 microcriths, of which 64 microcriths are oxygen. The oxygen is therefore $\frac{64}{98}$ of the whole weight, and $\frac{64}{98} \times 100 = 65.31$ lbs., which is the required answer.

Problems may be solved by proportion as follows: —

Molecular weight : microcriths of element
= weight compound : weight of element.

In above example we have: —

$$98 : 64 = 100 : x,$$
$$x = 65.3 \text{ lbs.} \quad \textit{Ans.}$$

Rule. — *Determine the number of microcriths in the molecule of the compound. Multiply the given weight of the compound by the number of microcriths of the element divided by the total number of microcriths in the molecule, and the result is the quantity of the element in the given weight of the compound.*

Case II

87. To compute the Weight of a Substance Required to Combine with, or Obtainable from, a Given Weight of Another Substance. — If substances always combined molecule for molecule, the weight of each substance required would be proportional to the molecular weights; but as this is the exception rather than the rule, it is necessary to consider the reaction to determine the number of molecules of each substance required. In the preparation of hydrogen, the following reaction occurs: —

$$Zn + 2\,HCl = ZnCl_2 + 2\,H.$$

It will be observed that two molecules of hydrochloric acid are required to form one molecule of zinc chlorid. Now one molecule of hydrochloric acid weighs 36.5 microcriths, and two molecules 73 microcriths. One molecule of zinc chlorid weighs $65 + 35.5 + 35.5 = 136$ microcriths. Therefore 73 microcriths of hydrochloric acid will form 136 microcriths of zinc chlorid, and larger weights of the two substances will be in the same proportion; hence the following: —

Rule. — *Write the reaction. Determine the number of microcriths of each substance involved in the problem. The ratio of the microcriths will equal the ratio of the actual weights.*

Example. — 1. What weight of hydrochloric acid can be made from 100 lbs. of sodium chlorid?

Reaction: $2\ NaCl + H_2SO = Na_2SO_4 + 2\ HCl.$

2 molecules of NaCl weigh 117 microcriths.

2 molecules of HCl weigh 73 microcriths.

$$117 : 73 = 100 : x,$$
$$117\,x = 7300,$$
$$x = 62.3 \text{ lbs. } \quad \textit{Ans.}$$

2. How much oxygen required to burn 10 grammes of carbon?

$$C + 2\ O = CO_2,$$
$$12 : 32 = 10 : x,$$
$$12\,x = 320,$$
$$x = 26\tfrac{2}{3} \text{ grammes. } \quad \textit{Ans.}$$

Case III

88. To find the Weight of a Litre of Any Gas. — In physics we learn that the specific gravity of a substance is the ratio of the weight of the substance to that of a like volume of some standard. Two standards are in common use, the air and hydrogen, and to avoid confusion it is customary to apply the term *vapor density* to the weight of a gas as com-

pared with hydrogen. The vapor density of a gas, therefore, simply expresses the fact that any volume of the gas is just so many times as heavy as a like volume of hydrogen under like conditions of temperature and pressure. The letters N.T.P. (normal temperature and pressure) are often used in connection with problems of this kind. One litre of hydrogen, N.T.P., weighs .0896 gramme. This quantity is called a *crith*.

The vapor density of each element thus far considered is equal to its *atomic* weight. The vapor density of *any* gas is equal to one-half its *molecular* weight.

Example. — To find the weight of 10 litres of oxygen.

From the table on p. 14 we find that the atomic weight of oxygen is 16. As it is an element, this number is also its vapor density; hence 10 litres will weigh .0896 gramme $\times$ 10 $\times$ 16 $=$ 14.33 grammes.

Ans.

Rule. — *Multiply the weight of one litre of hydrogen by the vapor density of the gas.*

Case IV

89. To compute the Percentage Composition of Any Substance. — Express as a decimal the ratio of the weight of each element to the weight of the compound. These decimals give the amount of each element in a unit weight of the compound, and in 100 weights there would be 100 times as much of each element. Move the decimal point, in each case, two places to the right, and the percentage composition is therefore obtained.

Note. — A litre of air weighs 14.44 criths.

REVIEW QUESTIONS

1. How many kilos of sulfur are contained in 49 kilos of sulfuric acid, H_2SO_4?

2. How many grammes of carbon dioxid, CO_2, should be obtained by burning a diamond weighing $\frac{1}{2}$ gramme?

3. Find the weight of oxygen in 100 grammes of ferric oxid, Fe_2O_3.

4. How much hydrochloric acid, HCl, can be made from 100 lbs. of salt, $NaCl$?

5. If 100 grammes of oxygen are to be converted into water, how many grammes of hydrogen must be used?

6. How many kilos of hydrochloric acid may be obtained from 20 kilos of salt?

7. How much potassium chlorate, $KClO_3$, would be required to evolve 2 lbs. of oxygen?

8. In 126 lbs. of nitric acid, HNO_3, how much oxygen?

9. Find the proportion of each element in chloric acid, $HClO_3$, by volume. By weight.

10. How many pounds of hydrochloric acid, HCl, can be made from the chlorin found in 234 lbs. of sodium chlorid, $NaCl$?

11. How many kilos of sodium chlorid would be required to prepare 73 kilos of hydrochloric acid?

Reaction: $2\ NaCl + H_2SO_4 = Na_2SO_4 + 2\ HCl.$

12. How many pounds of sulfuric acid will be required to convert 1010 lbs. of potassium nitrate, KNO_3, into nitric acid, HNO_3?

Reaction: $2\ KNO_3 + H_2SO_4 = K_2SO_4 + 2\ HNO_3.$

13. Find the number of grammes of sulfuric acid required to furnish 100 grammes of hydrogen, by its action on iron.

14. If 1 litre of oxygen weighs $1\frac{1}{4}$ grammes, how many litres of oxygen will result from the decomposition of 1 kilogramme of water?

15. If a liter of hydrogen weighs .089578 gramme and a litre of oxygen weighs 1.429 grammes, what is their proportion by weight in water?

16. If 806.4 grammes of water are decomposed, what is the volume of each of the resulting elements?

17. Find the number of grammes of potassium chlorate required to furnish 100 litres of oxygen.

18. Find the number of litres of hydrogen that may be set free by the action of sulfuric acid on 10 grammes of zinc.

19. Find the number of litres of oxygen obtained from 100 grammes of potassium chlorate heated with manganese dioxid.

20. Compute the percentage composition of potassium chlorate. How many grammes of oxygen would be given off by heating 500 grammes of the chlorate, and what would be its volume?

21. Find the percentage composition of water by weight and by volume.

22. A litre of water weighs 1000 grammes; how many grammes of hydrogen does it contain? how many grammes of oxygen? how many litres of each?

23. How many litres of the element in the first column may be obtained from 10 grammes of the substance in the second column?

H	H_3SO_4
O	H_2SO_4
O	HNO_3
H	HNO_3
N	HNO_3
H	HCl
Cl	HCl

24. How many grammes of the substance in the first column will be required to prepare 10 litres of the element in the second column?

H_2SO_4	H
$KClO_3$	Cl
$KClO_3$	O
Zn	H
H_2O	O

CHAPTER X

COMPOUNDS OF NITROGEN AND HYDROGEN

NITROGEN combines with hydrogen to form three compounds; namely: —

Ammonia,	NH_3
Hydrazine,	N_3H_4
Hydrazoic acid,	N_3H

AMMONIA

FORMULA NH_3 — MOLECULAR WEIGHT 17

90. Occurrence. — Ammonia occurs in small quantities in the air, being formed under certain conditions by the decay of animal and vegetable substances.

The chief source of ammonia is the ammoniacal liquor of the gas works, which is the water through which the gas has been passed to remove the ammonia formed by the decomposition of the coal.

91. Preparation of Ammonia, NH_3.

Experiment XLIX. (For two students.)

NOTE. — Experiment 50 may be performed with this one if the students will arrange the three bottles with tubes required before beginning this experiment.

1. Mix 4 grammes of ammonium chlorid and 8 grammes of calcium hydrate on a piece of glass, adding a few drops of water. Place the mixture in a flask with a suitable delivery tube, add sufficient water to cover the mass, and apply heat.

2. Collect three bottles of the gas by upward displacement. Set them aside, mouth downward.

3. Hold the Bunsen burner flame in a stream of the gas. Can you ignite it? Does the gas burn while the burner is held in the stream? Is the gas combustible? Can its flame be said to have a color?

4. Connect the flask to the series of bottles described in Experiment 50, and proceed with that experiment.

5. When leisure permits, test the three bottles of the gas in such manner as to enable you to answer the following questions: —

Is it a supporter of combustion? Is it heavier or lighter than air? Is it soluble in water? Test with a piece of moist pink litmus paper. Note the odor. CAUTION.

When ammonia is prepared as above, the action is as follows: —

$$CaH_2O_2 + 2\,NH_4Cl = 2\,NH_3 + CaCl_2 + 2\,H_2O.$$

Ammonia is also formed when nitrogenous organic matter is heated out of contact with the air, as in the process of making illuminating gas by heating coal, or in the process of making animal charcoal.

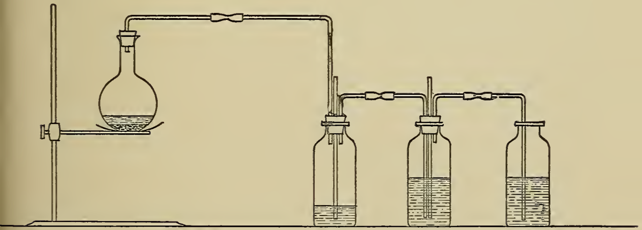

FIG. 8.

Ammonia gas is often prepared by heating the stronger ammonia water of commerce in a flask and collecting the gas as in Experiment 49.

Ammonia water was formerly called spirits of hartshorn, because it was prepared by distilling the horns of the hart.

92. Preparation of Ammonium Hydroxid, NH_4OH (Ammonia Water).

Experiment L. (For two students.) — 1. Connect a series of three medium sized bottles with the delivery tube of the flask used in Experiment 49 as shown in Fig. 8. Have 20 cc. of water in the first bottle, and about 50 cc. in each of the others. The first and second bottles are fitted with rubber stoppers. Neither tube of the first bottle should dip below the water. The tubes by which the gas enters the remaining bottles should dip beneath the water. The gas is sometimes dissolved faster than it is supplied in a given bottle, and a vacuum is formed, causing the water to run into it from the next bottle. Should this occur, raise the stopper of the bottle toward which the water flows. If three-holed stoppers are at hand, it is best to use them, and to insert safety tubes in the first and second bottles, thus preventing this action.

2. Hold a glass rod moistened with hydrochloric acid over a bottle of ammonia water. What occurs? This is the test for ammonia. Test the ammonia water with litmus paper.

93. Manufacturing Processes. — In the arts, ammonia is prepared either by adding slaked lime to the ammoniacal liquor of the gas works, or by boiling the liquor. In

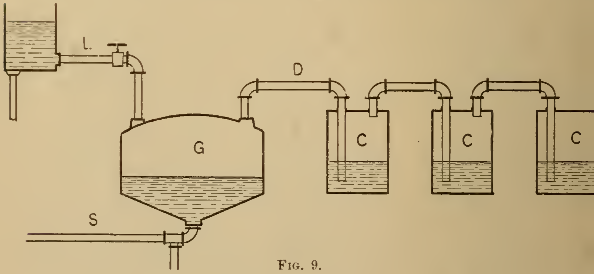

Fig. 9.

the latter process the vapor is often passed into sulfuric acid, forming crude ammonium sulphate, which is used as a fertilizer.

At the Rochester Ammonia Works the liquor is forced

into the generator *G*, Fig. 9, by a steam siphon *S*, and slaked lime is introduced through the pipe *L*. The mixture is agitated and heated, and the gas escapes through the delivery pipe *D* to the iron tanks *C*, where it is dissolved in water as in Experiment 50. Between *G* and *C*, the gas passes through several cylinders containing petroleum and various other solvents which absorb impurities, but allow the ammonia to pass through them. In this way, the "stronger ammonia water" of the drug stores is made.

94. Liquid Ammonia. — There is a large demand for anhydrous ammonia for use in ice machines. It is usually prepared as follows: Ammonia water is heated in an iron cylinder *A*, Fig. 10. The gas, NH_3, is driven over into the condenser *C*. As it accumulates, the pressure increases until it reaches 130 pounds per square inch, when the gas is liquefied.

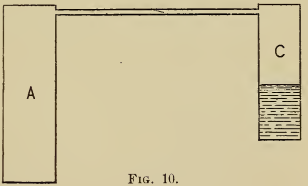

FIG. 10.

95. Physical Properties. — Ammonia gas can be easily condensed to the liquid form by cold and pressure. When the pressure is removed it passes back to the form of gas and absorbs heat in so doing.

One volume of water dissolves 600 volumes of ammonia gas at ordinary temperatures. It condenses at ordinary temperatures at 6.9 atmospheres; at ordinary pressures it boils at $-33.7°$ C. and solidifies at $-75°$.

96. Chemical Properties. — A jet of ammonia to which a flame is applied continues to burn in oxygen after the flame is withdrawn. In air, however, the heat of com-

bustion is not sufficient to raise adjoining particles to the kindling temperature.

In gaseous form as well as in the solution ammonia turns red litmus paper blue and neutralizes acids just as the alkalies, sodium, and potassium hydroxids do.

Ammonia combines directly with acids and many other substances to form a series of compounds which resemble each other in general properties, and each of which contains the group of atoms NH_4. In its chemical action, therefore, this group resembles a metallic element, and in order to distinguish it from ammonia, NH_3, the termination *um*, which is applied to nearly all metallic elements, is substituted for the final *a* of ammonia. The similarity in composition of some of the compounds of ammonium, zinc, and sodium is shown in the following table: —

NH_4	Zn	Na
NH_4Cl	$ZnCl_2$	$NaCl$
$(NH_4)_2SO_4$	$ZnSO_4$	Na_2SO_4
NH_4NO_3	$Zn(NO_3)_2$	$NaNO_3$

Compounds which play the part of an element are called *radicals.*

There is more or less evidence that a definite chemical compound is formed when ammonia is dissolved in water and that its formula is NH_4OH. This compound has never been isolated, however, and certain chemists doubt its existence. The name ammonium hydroxid implies that the solution is chemical rather than physical.

97. Uses. — Ammonia is extensively used in the manufacture of artificial ice, aniline colors, indigo, washing soda,

etc. It is also used in the laboratory as a reagent and in the household as a detergent.

The following device illustrates the manner in which liquid ammonia is employed in making artificial ice: —

Anhydrous ammonia flows from the tank *LA* into the chamber which surrounds a vessel of water *I*; it evaporates rapidly, and so much heat is rendered latent that the water is frozen. The gas is pumped back into the tank and again liquified by pressure, and the operation is repeated.

In the Carre Ice Machine, a pound of coal is consumed for each pound of ice made.

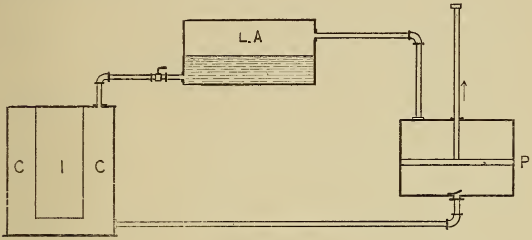

Fig. 11.

REVIEW QUESTIONS.

1. Describe the preparation of the ammonia of commerce.
2. Describe the preparation of ammonia in the laboratory.
3. Discuss the combination of ammonia with water.
4. Distinguish between ammonia and ammonium; between liquid ammonia and ammonia water.
5. In what respect does ammonium resemble a metal?
6. What is hartshorn?
7. State the color, odor, solubility, and weight of ammonia.